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INTRODUCTION  
TO THE  
STUDY OF CHEMISTRY  
PERKIN AND LEAN

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AN  
INTRODUCTION TO THE STUDY  
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AN

INTRODUCTION TO THE STUDY  
OF  
CHEMISTRY

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## PREFACE

THE assured position that Chemistry now holds in English education is shown by its inclusion as a compulsory or at least an optional subject in almost all examinations preliminary to entrance on the learned professions, and by the fact that many of the best schools now possess well-equipped laboratories and give regular instruction in the elements of both Theoretical and Practical Chemistry.

In directing the attention of our nation to the study of Chemistry as a means of scientific education, Sir Henry Roscoe has led the way. His text-books have given to generations of students their introduction to a knowledge of the science. His *Primer of Chemistry* and *List of Experiments* prepared for the Science and Art Department contain the subject-matter which every schoolboy has till recent times been expected to know, every examiner has asked for, and few text-books have dared to disregard. The output of admirable text-books on these lines has during recent years been so abundant, that the issue of another would be unjustifiable, if it differed from them merely in detail and in unimportant features.

But while the knowledge of the actual facts of Chemistry is worth much, there has been a growing feeling that the methods of gaining knowledge are often of more educational value than

the knowledge itself, and that a lad who starts life with his mind merely stored with facts is not so well equipped as one who has also *learnt how to learn*, not from books alone, but by direct contact with realities. In a school curriculum Science is chiefly valuable as a means of culture, and because there are faculties of the human intellect, namely, correct observation, sound reasoning from observation and experiment, and imagination in seeing into the invisible causes of visible phenomena, which can alone find their development in a carefully considered course in Physical Science. We are, therefore, even more anxious to insist on the methods by which knowledge has been obtained than on the actual facts of Chemistry. Sound methods of observation, moreover, remain valid, and may continually enlarge and correct our ideas in studying the further development of the science. At the same time the student must be sure of his facts before he begins to theorise. "Why is it," Charles II. is reported to have asked, "that when you put a dead fish into a bowl full of water it runs over, but if the fish is alive it does not run over?"

Some time ago Professor H. E. Armstrong, a leader in the movement which has sought to place the teaching of Chemistry upon what is regarded by many as a more rational basis, appended a detailed scheme for a school course on Chemistry to the Report upon Methods of Teaching Chemistry, which was presented to the British Association at Newcastle.<sup>1</sup>

Very recently, the Incorporated Association of Headmasters has issued a syllabus of a course of instruction in Elementary Science, including Physics and Chemistry, drawn up by a Committee on Science Teaching, of which Mr. C. M. Stuart, of Catford, was chairman, and Professor Armstrong a prominent member. To quote from this syllabus, "This course is intended for all

<sup>1</sup> See *Report of the British Association*, 1880. A subsequent paper elaborating certain features in this scheme may be found in *Nature*, vol. xliii. p. 593.

boys and girls commencing the study of Science. It represents in the opinion of the Committee a suitable commencement for those who continue the subject, and indicates the manner in which it may be made of true educational value to those who do not pursue it further. . . . While the main object of the course should be to train students to solve simple problems by experiment—to work accurately, and with a clearly defined purpose, and to reason from observation—the instruction given should eventually lead them to comprehend the nature of air, water, ‘fire,’ earth, and food.”

This syllabus of the Association of Headmasters reached us after we had nearly completed our work on the present book, and it was a matter of satisfaction to us to find that almost the whole of the syllabus and the detailed experiments for the course on Elementary Chemistry were dealt with in our manuscript.

While the aim of the course described in the present book, and the object of the courses referred to above have alike been the cultivation of a spirit of inquiry and the prosecution of Science as among the best means of culture, there is, nevertheless, one difference between the lines on which they have been worked out.

The principle which has been our general guide in the selection and treatment of subject-matter may be expressed in words used in another connection by Professor H. A. Miers, in a recent inaugural lecture<sup>1</sup> at Oxford, in which he applied the saying of the Darwinists that the development of the individual is an epitome of the evolution of the race.

“The order,” he said, “in which a subject can best be unfolded before a student’s mind is very satisfactorily marked out by the historical development of the subject : a profitable course of teaching is suggested by the history of a science, and *the order in which problems have presented themselves to successive*

<sup>1</sup> *Nature*, vol. liv. p. 208.

*generations is the order in which they may be most naturally presented to the individual."* This is, we believe, particularly applicable to the study of Chemistry on the part of beginners. At the same time the progress of Science has often of course been like a paper-chase with many false scents, and it is not necessary for the logical application of the above principle to follow all these blind alleys.

A glance at the Table of Contents will indicate the course which we have been led to adopt in our attempt to work out an Introduction to the Study of Chemistry. We have referred at the outset to Alchemy and to some of the errors which were current till the seventeenth century, showing the readiness with which errors arose unless checked by well-devised experiment and careful measurement. In this way the student is led to see the *necessity of practical work and of exact measurement at a very early stage*. Measurements are then made of length, of mass, of the volume of liquids, of temperature, of density, of pressure, and of heat.

Then follows the practice of important chemical operations, such as solution, filtration, evaporation, and crystallisation, and the student's powers of observation and description are exercised by the examination of certain important raw materials, found native in the crust of the earth. He next learns how to prepare the most important acids and alkalies from his raw materials, after the manner of Geber, Glauber and other alchemists.

Having thus acquired in some measure quickness of eye in observation, accuracy of expression in description, delicacy of hand in manipulation, and a passing acquaintance with the apparatus and chemicals in the laboratory, the student is ready to set to work upon a series of researches on the Action of Acids on Chalk, Fire and Air, the Rusting of Iron, the Discovery of Oxygen, the Action of Acids on Metals, Water, and other subjects, in which he is led to work so far as may be along the



lines which proved successful with Black, the Oxford chemists of the seventeenth century, Priestley, Scheele, Lavoisier, Cavendish, and other famous investigators. Many of the experiments introduced have been suggested by their researches, and frequent quotations of interest have been made from their accounts of their own experiments.

The student will thus have become acquainted with several Gases, and is therefore ready to investigate some of their principal properties. Then follow researches which have a direct bearing on the law of Definite Proportions ; for example, a complete research is made upon Chalk, following the lines of Black's classical investigation, and upon the quantitative composition of Water, and lastly, following Bergman, Richter, and others, the student is led to investigate the Equivalence of Acids and Alkalies, and of Metals. A chapter on the Law of Definite Proportions closes the book.

We hold then that not only the *order*, but also the *method*, by which Chemistry can best be unfolded before a student's mind, is suggested by its historical development. It is a necessary corollary that Chemistry cannot be learnt in the lecture-room alone, and that the *experiments should be for the most part quantitative in character*. Qualitative analysis has hitherto been largely used as the practical exercise in the science. It has some merits. We think, however, that practical chemistry should not consist, to such an extent as is usually the case, of "qualitative analysis," and that the right place for such a course is after and not before a simple quantitative course. While practical work should be introduced at a very early stage the lecture-room has of course a necessary place, and the student should not start an experiment until he thoroughly understands what the object of the experiment is, what he means to do, and how he means to do it.

There are somewhat rigid limitations of time and circumstance

which must be considered in selecting experiments suitable for a school laboratory. The experiments must not require more than two pairs of hands ; the manipulation must not be so difficult that accurate results cannot be obtained ; the apparatus must be cheap enough to allow of all the boys doing the same experiment at once, and the whole experiment must not take more than 1 or  $1\frac{1}{2}$  hours, including the time required for putting the apparatus away. These conditions exclude many important experiments ; for instance, the determination of the composition of water by passing hydrogen over copper oxide may be done by the teacher as a lecture table experiment, but it must generally be omitted by the class. Weighings should generally be conducted to the nearest centigram, because a boy's errors of manipulation are apt to exceed 1 centigram, and so milligram weighings will only give a fictitious accuracy and waste time. One per cent. accuracy is the standard that may be aimed at. We have dwelt at some length upon the details which are necessary to exactness in the results of experiments. It is a great advantage to have several boys working at the same time at an experiment, for example, on the weight of a litre of air or on the percentage of water in a crystal, and to let each repeat the experiment as often as possible ; results should afterwards be compared and discussed. A great deal of time is saved by letting boys work in pairs. Moreover, a spirit of co-operation is cultivated in this way. They learn at least as much as in doing the experiment alone, and perhaps more, because they teach one another.

If the separation of the mathematical from the practical treatment of the Metric System in Chapters II. to V. is criticised on *a priori* grounds as an error and an abandonment of our guiding principle, we would reply that the Metric System had not a gradual growth, but was the outcome of the labours of a Commission. It is futile to "discover" that 1 c.c. of water

weighs 1 gram. Moreover, we would add that experience shows that a student's mathematics should be ready before he touches the apparatus of the laboratory. And this applies not only here. If any would-be chemists have not yet mastered the elements of arithmetic, decimals, the unitary method, percentages and proportional parts, we recommend them to close this book and go back to their ciphering. There can be no sound knowledge of Physics or of Chemistry without mathematical backbone. There is nothing more distracting to teacher and to student than to find that laboratory results cannot be worked out for want of adequate mathematical knowledge.

It may perhaps also be objected that we have abandoned the principle we laid down as our guide in relegating Black's research on Chalk to so late a position. Our reply is that the aforesaid principle while invaluable as a general guide, must be applied with discrimination, and other considerations must not be lost sight of. For instance, it is desirable that a study of "everyday" phenomena of great interest, such as Fire and Air, should precede what is after all of a more special character. Moreover, the research on Chalk is the most complete in itself of any entered upon, and one which makes too great a demand upon the powers of hand and of mind to be introduced in its true chronological place immediately after the discovery of Fixed Air.

The value of this book would be little in practice, whatever its general merit, if it came straight from the writing-desk, and had not been employed in actual class work. Not only have many of the chapters been worked through by elementary students in the laboratories of the Owens College, but we have been most fortunate in securing the co-operation of Hugh Richardson, M.A., Science Master of Sedbergh School, and of S. H. Davies, M.Sc., Head of the Chemical Department of the Battersea Polytechnic, London. Using typed copies of our MSS., they have taught through almost the whole in their

classes. We have amended our text wherever their trial of it showed alteration to be required, and we are convinced that the practicability and success of our scheme in actual class work has thus been secured and verified. Mr. Richardson also made at our request the first draft of the chapters on the Properties of Gases, and he has from the first helped us with his experience in the teaching of Science in schools.

We gladly thank also Mr. S. M. Walford, of the Hyde and Oldham Technical Schools ; Mr. Foster, of the Nelson Technical School, and Mr. S. E. Brown, of the Friends' School, Ackworth, for testing some chapters in their own classes.

Our colleagues, Mr. G. J. Fowler, Dr. A. Harden, and Dr. W. A. Bone, have also given us kind assistance in certain details, while Mr. P. J. Hartog has helped us with criticism upon our chapter on the Law of Definite Proportions.<sup>1</sup>

We have pleasure in acknowledging in addition the services of many students in our laboratories in the working out of methods and experiments. In this respect we especially thank Messrs. Hibbert, Rose, Rowntree, Smith, and Whatmough, and Misses Fish and Scotson.

While we have spared no trouble in a task which has been no light one, we cannot hope that we have avoided all error either in judgment or in points of detail, and suggestions and corrections offered by fellow teachers will be freely welcomed.

We may add that it is our desire to include in any subsequent edition chapters, which are already in hand, on the Discovery of the Metals, on the Diffusion of Gases, and on Fuel and Food-Stuffs.

W. H. P.  
B. L.

OWENS COLLEGE,  
July, 1896.

<sup>1</sup> See *Nature*, vol. i. p. 149, and *British Association Report*, 1894, p. 618.

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## CHAPTER I

### THE BIRTH OF CHEMISTRY

**The Earliest Book on Chemistry.**—Chemistry is one of the most modern of sciences, and one which has been placed upon a sure foundation only in the course of the present century. Yet the actual beginning and the dawning knowledge of Chemistry can be traced back to Egypt in the early years of this era.

The oldest chemical record<sup>1</sup> of which we have any knowledge is a papyrus found at Thebes in the wrappings of the embalmed body of a goldsmith, who had lived about the time of the second or third century after Christ. From this note-book of the goldsmith, which is now in the library of the University of Leyden, we find that he was acquainted not only with gold and silver, but also with copper, lead, tin, zinc, and arsenic. Recipes are given for preparing imitation gold and silver, and for increasing the weight of gold and silver by the addition of inferior metals such as lead without altering the apparent character of the former.

The notes of this fraudulent workman, which have been preserved in perfect condition in an Egyptian mummy, and have escaped the accidents of fifteen hundred years, show that at that time the knowledge of the metals possessed by some of the Egyptians was very considerable. Not only could the author counterfeit true gold or silver, the "noble" metals as they were called, by mixing inferior or "base" metals, but

<sup>1</sup> See Berthelot, *Les Origines de l'Alchimie*.

tests are given by which the imitation could be detected and distinguished from the pure and genuine metal.

For instance, it is stated that gold should keep its colour when melted. If it becomes whitish it contains silver, if it blackens it contains lead, and if it becomes rough or hard when cold it contains copper or tin. So also silver should remain white and shining when fused. If it is blackened it must be falsified with lead, and if it is rendered hard and yellow, copper is present.

**Alchemy: The Transmutation of Metals.**—The imitation of the noble metals by the process of mixing the base metals in various proportions, as the writer of the papyrus described in the fourth century, was accepted in the next century as an actual change or *transmutation* of the base metals into the noble metals. Extraordinary as it may seem to us, for more than a thousand years such transmutation was believed in not only by the ignorant and unlearned but even by philosophers.

Indeed, it seems probable that the very word Chemistry is derived from the Greek *Chemeia* or *χημεία*, transmutation, since the transmutation of the metals was the chief object of the Egyptians.

About the year 640 the Arabians overran Egypt, and becoming acquainted there with Chemistry, they prefixed to the word for it the Arabic Article *al*, so that with them the science was spoken of as *Alchemy*. The Arabians spread through Northern Africa into Spain, and founded many Universities to which students flocked from all parts of Europe.

From the Egyptians the Arabian Alchemists had acquired the firm belief in the possibility of the change of one kind of material into another, and of the transmutation of the base into the noble metals. Their most earnest endeavours were given to the search for the *philosopher's stone*, a substance of such extraordinary virtue that according to Roger Bacon one part of it was believed to have the power to convert 1,000,000 parts of base metal into pure gold. The possibility of the existence of such a substance was accepted by the most learned; for instance, Van Helmont, a distinguished chemist, a physician and an honest man, recorded how in 1615 he obtained one quarter of a grain of the philosopher's stone and converted eight ounces of mercury into gold. Helvetius, again, was an opponent of the Alchemists until 1666, when he managed to obtain a little piece of the philosopher's stone with which he said he

converted lead into a yellow metal, afterwards tested by the Master of the Mint and pronounced to be gold.

In some cases we can trace the fraudulent means by which the apparent transmutation was effected. A certain Duke of Wurtemberg kept an Alchemist in his service, and following his directions, the Duke was accustomed to place the necessary ingredients in a crucible, and then, all having left the room, the door was locked, and the next morning, behold, pure gold was found in the crucible. It is recorded, however, that the gold was obtained, not by the miraculous agency of any philosopher's stone, but by the hand of the Alchemist's son, who had been concealed in a cupboard within the laboratory.

Remarkable as the belief in the transmutation of the metals appears to us, many facts and experiments may be mentioned which at first sight seem to confirm such a belief. The following experiments will illustrate this point:—

**EXPT. I. To Obtain Lead from Galena.**—Galena is a metallic-looking mineral. Grind a small portion to a

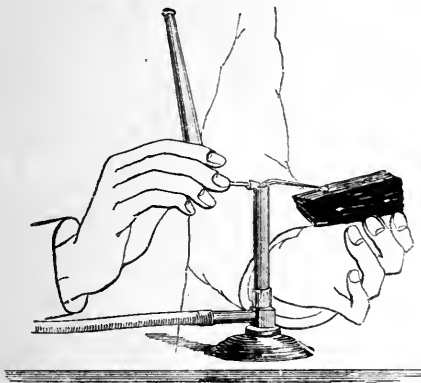


FIG. 1.

fine powder in a mortar. Make a small cavity in a lump of charcoal, and place in it a little of the powder (as much as a small shot). Turn a Bunsen flame so as to form a small luminous flame. Place the nozzle of the mouth-blowpipe just within the flame, and force a gentle current of air into it.

Direct the tip of the flame upon the galena on the charcoal. Heat till white fumes are evolved, and a bright mirror-like surface of molten metal appears, which will perhaps gather into a bead. Allow it to solidify, and then transfer the bead to a mortar and crush it. It is flattened out into a little plate of metal. It is soft, marks paper, and has all the characteristics of lead.

The production of lead from galena was familiar to the Alchemists. Moreover, since by suitable treatment they found how to extract a small amount of silver from a large amount of lead, they supposed that an actual change of galena into lead and then of lead into silver was effected. It does not seem ever to have occurred to the Alchemists that the galena at the outset perhaps contained lead and silver. Their error, for it is now known to be such, might have been avoided if they had paid any attention to the weights of the substances employed and formed in their experiments.

**EXPT. 2. To Convert Zinc apparently into Copper.**

—Pour some solution of blue vitriol (copper sulphate) into a test-tube till it is one-third full. Cut some strips of zinc narrow enough to enter the tube, drop in one strip, and notice the change of appearance. Boil, and observe that the zinc disappears, while a red and black deposit of copper is formed, as though the zinc had been transmuted into copper. Pour the solution into another test-tube, add another strip of zinc, and see whether any more copper is deposited on boiling.

**EXPT. 3. To Convert Copper apparently into Silver.**—Clean a strip of copper thoroughly by means of sandpaper. Place it for a moment in a solution of sulphate of mercury.<sup>1</sup> On removing the metal from the solution it is found to have acquired a bright white silvery surface, as though the copper had been transmuted into silver. Rub the strip with sandpaper, and the bright silvery appearance is destroyed. The copper therefore has not at any rate been entirely changed into another metal, although, at first sight, such would appear to have been the case.

These experiments show that first impressions and ideas as to phenomena may be very far from the truth, and should

<sup>1</sup> Nitrate of mercury may be used in place of the sulphate, though the result is not quite so good.



in every case be severely tested and supported by further experiment before they are finally accepted.

**Conclusions must be tested by experiment and by measurement.**—Another illustration of the way in which ideas may change is found in the opinions which have been held as to whether dead matter of itself can ever become alive.

The Roman poet Vergil<sup>1</sup> records the notion as being widely prevalent in Egypt, Persia, and the East, that bees could be generated from the putrid carcasses of oxen. He specially recommends a two-year-old bullock for the purpose, and prescribes the mode of killing it by suffocation and beating, then the pounding and mashing of it, hide and all, and laying it on thyme and cassia in a small shady chamber, with access of air at fit intervals, and so forth. Vergil does not tell us that he had seen it done, nor does he hint that the occasional selection by a swarm of bees of the hollow of a beast's skeleton as a convenient home may have originated the idea on which he enlarges with so much minuteness.

Our forefathers certainly believed that dead matter could beget life, and would have instanced vinegar, which breeds vinegar-eels. Yet this great question of spontaneous generation has since been answered by experiments, which show that so far as we know matter never becomes alive except by means of pre-existing life.

The laborious and life-long studies of the Alchemists did much to increase the knowledge of chemical operations, of minerals and metals, and of chemical agents, but the true method of scientific investigation was not theirs. They were too ready to be satisfied with a shallow explanation of what they saw. They were not accustomed to weigh the materials with which they began an experiment, nor to determine the weight of anything left at the end. In fact measurement was neglected by them. The result was that they were led to conclusions which were to a large extent erroneous. Nature does not reveal her secrets to the careless, nor to impatient enquirers. And until the chemist made the balance his constant servant, and was willing to test all his conclusions by rigorous experiment, there was and could be no sure foundation for the theories and laws which he endeavoured to construct.

<sup>1</sup> Vergil, *Georgics IV.*, 281-314.

## CHAPTER II

### THE METRIC SYSTEM

**Standards of Measurement.**—It has already been insisted upon that measurement is the first and most important step in the progress of any science, and it will be observed that in the statement of any measurement there is the mention firstly of a number, and secondly of a thing of the same kind as the quantity to be measured, which is referred to as a *standard* or *unit*. For instance, if it is desired to state the sun's distance from the earth, the distance of the moon from the earth may be selected as a standard or unit, and it may be stated that the sun's distance is 400 moon's distances. Similarly, the moon's distance may be stated in terms of the earth's radius, as about 60 earth's radii. But whilst such statements convey definite meanings to the mind, since in either case the magnitude of the unit chosen is comparable with the quantity to be measured, yet it would be very inconvenient in scientific work to state quantities in terms of standards not universally recognised, for it would be extremely troublesome to compare the statement of one man with that of another if they employed different standards. It is on this account that a statement of quantities in yards, gallons, or pounds is almost meaningless to a Frenchman or German. On the other hand the "Metric System" of measurement, to which reference is made below, is understood by scientific men of all nationalities, and indeed is exclusively employed in scientific investigations. There is another objection to the English system for purposes of measure

ment in the fact that there is no simple relation between the different standards. There is consequently much trouble in making calculations with quantities expressed by means of them. The Metric System, on the other hand, is a decimal system, and the different standards are related in a simple way to one another.

We shall now refer to the standards of Length, Area, Volume, and Weight in the Metric System, which are used in chemical investigations, and explain the nomenclature used in connection with this system.

### Metric System

**Standard of Length.**—The **Metre**<sup>1</sup> is the length at 0° C. of a certain bar of platinum kept in Paris.

In 1790, immediately after the French Revolution, a commission of savants nominated by the French Academy was appointed to prepare a new system of units.

Borda, Lagrange, Laplace, Monge, and Condorcet, who formed this commission, decided in favour of a decimal system, with the unit of length, from which all other units should be derived, itself connected with the size of the earth. This happy idea put the unit beyond reach of change from temperature or from other causes.

The commission added to itself Lavoisier and other eminent men, and had a meridian arc from Dunkirk to Barcelona carefully measured. The metre was intended to be  $\frac{1}{10,000,000}$  of the meridian quadrant of the earth's circumference drawn through Paris (Fig. 2). This would make the circumference of the earth 40,000,000 metres. A platinum metre of the required size was constructed. Since then a small error has been found

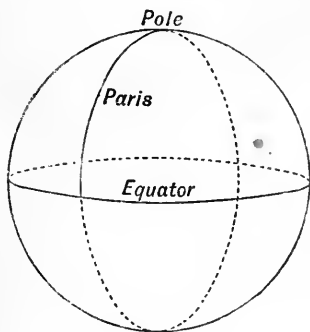


FIG. 2.

<sup>1</sup> Greek: *metron*, a measure.

in the calculations, so that this platinum metre is not exactly what it was intended to be. But Borda's metre is still taken as the ultimate standard of appeal. It was kept for some time in the Musée des Archives in Paris, but has since been removed to the Ministère de l'Intérieure. Careful copies have been made and distributed among the European nations.

The metre is divided decimally into tenths, hundredths, and thousandths. Names are given to each of these divisions. Other units are equal to 10, 100, and 1,000 times the metre.

The *prefixes* denoting multiples of metres are derived from the Greek numerals *χίλιοι*, *έκατον*, *δεκα*, those denoting decimal parts of the metre from the Latin *decem*, *centum*, *mille*. The same prefixes are used in the tables of capacity and weight. Thus we have :—

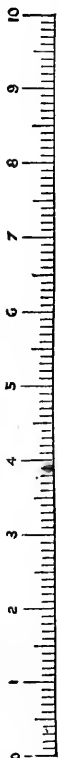
1 kilometre	equals	1000 metres	—denoted by	Km.
1 hectometre	„	100	„	Hm.
1 dekametre	„	10	„	Dm.
1 metre			„	m.
1 decimetre	„	·1	„	dm.
1 centimetre	„	·01	„	cm.
1 millimetre	„	·001	„	mm.

Figure 3 represents the exact size of one decimetre or one-tenth of a metre, divided into 10 centimetres ; and each centimetre is divided into 10 millimetres.

The student should try to remember roughly these metric measures by reference to real objects of known size. Thus :—

Kilometre	equals	distance covered by 10 minutes' quick walking.
Hectometre	„	length of football field.
Dekametre	„	half length of cricket pitch.
Metre	„	width of door.
Decimetre	„	width of hand.
Centimetre	„	width of finger-nail or dice.
Millimetre	„	nib of J pen, or thickness of sixpence.

FIG. 3.



## EXAMPLES I

## Metric Prefixes—Length

1. How many millimetres make 1 kilometre ?
2. Express 1 centimetre as a decimal of a decimetre.
3. What fraction is a decimetre of a dekametre ?
4. Express 1 millimetre in kilometres.
5. Subtract 1 millimetre from 1 centimetre ; answer in centimetres.
6. Add 1 metre to 1 Km. ; answer in Km.
7. What unit is equal to 1000 cm. ?
8. Express 4'56 Hm. in cm.
9. Bring 12'3 dm. to Dm.
10. How many cm. are there in 123'456 metres ?
11. The diameter of a halfpenny is 25 mm. How many halfpennies in a row would stretch 1 metre ?
12. How many metre strides shall I take in running 5 kilometres ?
13. If a sixpence is 1 mm. thick, find the height in cm. of a pile of sixpences worth £1.
14. If a cricket pitch is 2 dekametres long, how many kilometres does a boy run in making a century ?
15. A fishing rod has 3 joints and a total length of 3'6 metres. Find the average length of each joint in cm.
16. How many hours will it take to walk 24 kilometres at the rate of 1 hectometre per minute ?
17. How many pins, each 25 mm. long, can be made from 1 kilometre of wire ?
18. Four pieces of 1 decimetre each are cut away from a metre of glass tubing. Into how many lengths of 15 cm. long can the remainder be divided ?

**Standard of Area.**—The standard of area is the square metre represented by the symbol sq. m. The multiples and sub-multiples of this are named as before (p. 8), and the table on that page refers to this case also, if the prefix sq. is always put before the quantities.

Since 1 decimetre = 10 centimetres,  
therefore, 1 square decimetre =  $10 \times 10$  or 100 square centimetres.  
Thus,

$$1 \text{ sq. cm.} = 100 \text{ sq. mm.}$$

$$1 \text{ sq. dm.} = 100 \text{ sq. cm.}$$

$$1 \text{ sq. metre} = 100 \text{ sq. dm.}$$

In order to obtain a definite idea of the size of the sq. decimetre, sq. centimetre, or sq. millimetre, refer to the following figure :—

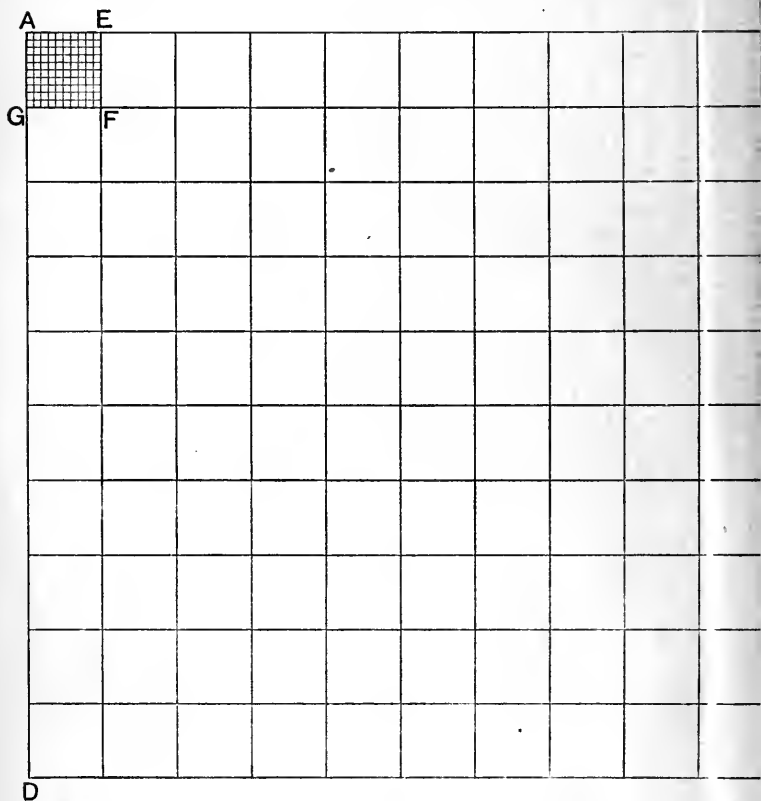


FIG. 4.

ABCD is 1 square decimetre. It is divided into 100 sq. cm., one of which AEFG is divided into 100 sq. mm.

## EXAMPLES II

**Metric Square Measure—Area**

1. How many sq. cm. are there in 1 sq. metre?
2. Express 1 sq. mm. as a decimal of 1 sq. cm.
3. If 1 inch equals 25 mm., how many sq. mm. equal 1 sq. in.?
4. If 1 sq. foot equals 900 sq. cm. how many cm. equal 1 linear foot?
5. Add 1 sq. m. to 2 sq. dm., and express in sq. cm.
6. Subtract 4 sq. mm. from 3 sq. cm., and reduce to sq. metres.
7. Find 80 mm.  $\times$  25 mm., in sq. cm.
8. Multiply 3 dm. by 2 mm.; answer in sq. cm.
9. Divide 16 sq. cm. by 8 mm.; answer in cm. long.
10. Find in sq. dm. the area of a post-card 12.5 cm. long and 80 mm. broad.
11. If 1 dm. equals 4 ins. long, express 1 sq. metre in sq. yds., sq. ft., sq. ins.
12. If the pressure of the air is  $2\frac{1}{8}$  lbs. per sq. cm., what is that in cwts., lbs., per sq. metre?
13. How many postage stamps each 24 mm. by 20 mm. are required to paper a wall 4 metres by 3 metres?
14. How many pieces of zinc each 25 cm. by 8 mm. can be cut from a sheet 16 dm. by 8 cm.?

**Standard of Cubic Contents or Volume.**—A cubic metre is a cube, each edge of which is 1 metre long. This volume is represented by cu.m., and it is subdivided in the usual way into cubic decimetres (cu.dm.) and cubic centimetres (c.c.). The cubic centimetre is the standard of volume in laboratory work.

Since contents = length  $\times$  breadth  $\times$  depth, therefore 1 cubic decimetre =  $10 \times 10 \times 10$  or 1000 cubic centimetres, and

$$1 \text{ c.c.} = 1000 \text{ cu. mm.}$$

$$1 \text{ cu. dm.} = 1000 \text{ c.c.}$$

$$1 \text{ cubic metre} = 1000 \text{ cu.dm.}$$

**Standard of Capacity.**—Gases and liquids which take the shapes of the vessels containing them may be measured by cubic measure. They are generally measured in cubic units, but under other names. It has been found convenient in measuring liquids to have a distinct standard.

This standard is *One Litre*. It is the volume of one cubic decimetre.

The multiples and decimals of a litre are named like those of the metre ; thus a decilitre equals  $\frac{1}{10}$  litre.

Since 1 litre = 1 cu. decim. = 1000 c.c.,  
therefore, 1 decilitre =  $\frac{1}{10}$  litre = 100 c.c.

Thus — 1 litre = 1 litre = 1000 c.c.

1 decilitre =  $\frac{1}{10}$  „ = 100 c.c.

1 centilitre =  $\frac{1}{100}$  „ = 10 c.c.

1 millilitre =  $\frac{1}{1000}$  „ = 1 c.c.

It is important to have some idea of the real sizes of these measures. Ask your teacher to place before you the following objects, possessing about the capacities stated :—

Large flask of capacity	. . . . .	1 litre.
Tumbler	„ „ . . . . .	$\frac{1}{4}$ litre.
Wine-glass	„ „ . . . . .	1 decilitre.
Pipette	„ „ . . . . .	1 centilitre.
Die	„ „ . . . . .	1 millilitre.

### EXAMPLES III

#### Metric Cubic Measure

1. If 30 cm. equal 1 foot, how many c.c. make 1 cu. ft.?
2. If 25 mm. equal 1 inch, how many cu. mm. make 1 cu. in.?
3. If 16 sq. ins. equal 1 sq. dm., how many cu. ins. make 1 cu. decim.?
4. How many cu. mm. are there in 1 cu. dm.?
5. Express 1 cu. mm. as a decimal of 1 c.c.
6. How many cu. mm. make 1 cu. metre?
7. Express 1 c.c. in cu. metres.
8. Add 1 c.c. to 2 cu. mm., and express in cu. mm.
9. Express in c.c, 3 cu. metres + 4 cu dm.
10. Subtract 6 cu. mm. from 5 c.c., and express the answer in c.c.
11. A cubical die measures 1.1 cm. each way. What is its volume in cu. mm.?
12. The lead of a pencil is 2 mm. broad, 2 mm. thick and 15 cm. long, what is its volume in cu. mm.?
13. How many cu. dm. equal 1 cu. ft. (30 cm. equal 1 ft.)



14. Find in c.c. the contents of a biscuit box 1.5 dm. high, 12 cm. broad and 240 mm. long.

15. How many cu. dm. of water are contained in a bath 60 ft. long, 24 ft. broad and 5 ft. deep? (See question 13.)

16. How many c.c. of mercury can be contained in a tube 1 sq. mm. in section and 1 metre long?

17. A measuring jar contains 250 c.c. of water, and is 25 cm. high. Find the area of its cross section.

#### EXAMPLES IV

##### Metric Capacity Measure

1. Express in c.c. 1 litre + 2 decilitres.
2. Bring 3 centilitres + 4 millilitres to c.c.
3. How many millilitres are equal to 5 dl. + 7 cl.?
4. Express 6 ml. + 8 dl. in centilitres.
5. Subtract 2 dl. from 1 l. and express the answer in c.c.
6. What decimal of 1 cu. dm. is 1 decilitre?
7. Divide 1 decilitre by 1 c.c.
8. Express 1234 c.c. in litres.
9. How many c.c. are there in 56.78 litres?
10. How many decilitres are there in  $\frac{1}{100}$  of a cubic decimetre?
11. Express (1 decilitre +  $\frac{1}{100}$  cu. decimetre + 100 c.c.) in litres.
12. Express (1 cu. decimetre + 10 decilitres + 1000 millilitres) in c.c.
13. A solution of salt is to be diluted ten times. How many centilitre pipettes full of the salt solution must be taken to make 500 c.c. of the dilute solution?
14. Find in litres the contents of a pneumatic trough 80 mm. deep, 25 cm. long and 1 dm. broad.
15. How many times can a dekalitre bucket be filled from a trough 2 metres long, 50 dm. broad and 40 cm. deep?
16. 600 cu. ft. of air each is allowed in a dormitory for 20 people. How many litres is that (1 foot long equals 30 cm.)?
17. How many days will a hectolitre cask of water last for 7 people, allowing 3 pints per day (1 litre equals  $1\frac{3}{4}$  pints)?

**Standard of Weight.**—In England the relations between the measures of the volumes and weights of given quantities of water are complicated. In the Metric System the relations are simpler.

*One Gram* (gram.) is the weight of 1 cubic centimetre of water

at 4° C. Multiples and decimals of 1 gram are named as before. Thus,

1 kilogram	=	1000 gm.
1 hectogram	=	100 gm.
1 dekagram	=	10 gm.
1 gram		
1 decigram	=	·1 gm.
1 centigram	=	·01 gm.
1 milligram	=	·001 gm.

The Dekagram may be remembered as the weight of the French 2 franc piece; it is slightly less than the weight of a florin, and more than that of a penny. A threepenny-piece weighs about  $1\frac{1}{4}$  grams.

#### EXAMPLES V

##### Metric System—Weight and Volume

1. A litre flask weighs 150 grams when empty; how many grams will it weigh when full of water?
2. A decilitre flask weighs 125 grams when full of water. How many grams will it weigh when empty?
3. A biscuit tin measures 16 cm., by 25 cm., by 20 cm. How many kilograms of water will it hold?
4. A solution of salt is to be made up so that the salt is 5 per cent of the weight of the water. How much salt must be taken to a litre of water?
5. Mercury is  $13\frac{1}{2}$  times as heavy as water. How many grams of mercury will fill a decilitre flask?
6. Sulphuric acid is 1·8 times as heavy as water. How many cc. of sulphuric acid will weigh 99 grams?

##### The Relations Between English and Metric Units:—

Some rough equivalents are worth remembering.

1 foot	equals	30 cm.
1 inch	„	25 mm.
1 decimetre	„	4 in.
1 kilometre	„	$\frac{5}{8}$ mile.
1 ounce	„	28 grams.
1 litre	„	$1\frac{3}{4}$ pints.

More exact numbers are useful for reference :—

1 metre	equals	39·37 inches.
1 foot	,,	30·5 cm.
1 gram	,,	15·43 grains.
1 kilogram	,,	2·20 lbs.
1 litre	,,	1·76 pints.
1 litre	,,	61·03 cu. ins.

*Note.*—If 1 gram equals 15·43 grains, then 1 grain equals  $\frac{1}{15·43}$  gram, or ·06489 gram.

The numbers 15·43 and ·06489 are called *reciprocals*.

**Some Disadvantages of the Metric System.**—The advantages of the Metric System over any existing system for the purposes of scientific investigation are great. It is understood by men of all nationalities, and the decimal relations which exist between the units are of great convenience in calculations.

It is not, however, urged here that the Metric System is the best for all the ordinary transactions of life. Indeed, as Herbert Spencer has recently pointed out, there is some ground for the belief that a decimal system is not the most convenient for the purposes of trade. Quoting from a letter which appeared in the *Times*, he gives the following instance : “ In England, if an old village woman buys a measured quantity involving fractions of an article at a price reckoned, say, in shillings, pence, and halfpence, the shopman or girl never seems to find any difficulty in working quickly in his or her head the sum which has to be paid, and the old woman manages to see that she is charged correctly. But, in France, with the decimal system, it seems always to be necessary to work the simplest sum out on a piece of paper. In fact, it is not difficult to see that it is pretty easy to work out mentally the value of, say, a yard and three-quarters of riband at  $3\frac{1}{2}d.$  a yard, while it is very difficult to work out mentally the value of, say, 1 m. 75 centimetres of riband at 35 centimes a metre.”

Moreover, in different trades different units are found to be convenient, and though the Government of France enforced as long ago as 1839 the observance of the metric system, it is remarkable that its use is even to-day not general in France. “ Precious stones are to-day bought and sold in carats ; firewood

in cordes ; milk in pintes ; gravel in toises ; grain, potatoes, and charcoal in boisseaux ; sugar and tea among the poor people is dealt with in livres, demielivres, &c."

### EXAMPLES VI

#### Metric System—Miscellaneous

1. How many pieces of glass rod each 5 cm. long can be cut from a piece 1 metre long ?
2. How many times can a 10 c.c. pipette be filled from a half-litre flask ?
3. If 1 metre of magnesium ribbon weighs 4 decigrams, how many milligrams will 5 cm. weigh ?
4. A flask weighs 20 grams when empty and 120 grams when full of water. What will it weigh when full of mercury of which 1 c.c. weighs 13.6 grams ?
5. If a filter paper 20 centimetres square leaves 4 milligrams of ash when it is burnt, how many grams of ash will be left by 20 square centimetres of the same paper ?
6. If a platinum decigram weight measures 1 sq. cm., find in grams the weight of a piece of platinum of the same thickness measuring 4 mm. by  $2\frac{1}{2}$  mm.
7. A litre of salt solution contains 60 grams of salt. How many c.c. of water must be added to make the solution of the same strength as one containing 40 grams of salt per litre ?
8. A train is travelling 100 kilometres per hour. What is that in centimetres per second ?
9. If 15.43 grains equal 1 gram, express 1 grain in grams to 4 places of decimals.
10. Find in £ s. d. the cost of a journey from Paris to Boulogne—160 miles—at 15 centimes per kilometre. (10 centimes equal 1 franc ; 25 francs equal £1 ; 8 kilometres equal 5 miles.)

## CHAPTER III

### THE MEASUREMENT OF LENGTH

WE have already stated that the Alchemists and all students of Chemistry up to the seventeenth century gave little or no attention to measurement in their experiments. This neglect of measurement led to many false ideas, which were only slowly and with difficulty dispelled before the fundamental principles of the science of Chemistry, as it is now known, could be established.

The foundations of Chemistry rest upon experiment, and all exact experiment involves measurement. It is therefore necessary for us to become acquainted at this stage with the most important methods of measurement.

#### The Measurement of Length

**The Use of a Scale.**—Obtain a small box-wood rule marked with both a centimetre scale and an inch scale. Examine the graduations.

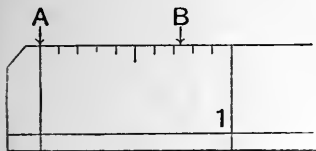


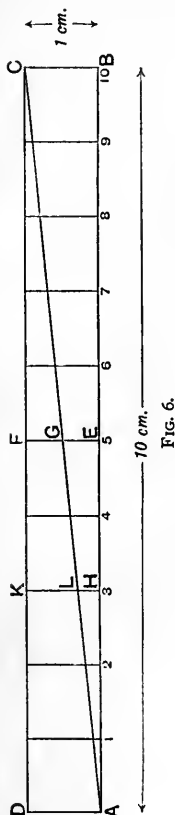
FIG. 5.

If an inch scale divided into tenths is available it is worth while to compare it with a centimetre scale. Observe that it is possible to estimate

lengths by means of such a scale graduated in tenths of an inch, to the nearest  $\frac{1}{100}$  inch. We may imagine each  $\frac{1}{10}$  inch to be divided into 10 smaller but invisible divisions of  $\frac{1}{100}$  inch each.

Figure 5 represents the edge of an inch scale.

AB =  $\cdot 7$  inch to the nearest  $\frac{1}{10}$  inch.  
 =  $\cdot 7$  + inch, *i.e.* between  $\cdot 7$  and  $\cdot 8$  inch.  
 =  $\cdot 73$  inch, estimating hundredths of a division.



## EXERCISES

### (i) Measurement of Length

1. Place a 10 to the inch scale against a scale divided into centimetres and millimetres. Read off in inches and decimals the length of 1 cm., 2 cm., 3 cm., . . . . . 10 cm., estimating to the nearest hundredth of an inch. For instance,

1 cm. =  $\cdot 4$  inch.

2 cm. =  $\cdot 79$  ,,

2. Place two scales together as before. Read off in millimetres (and by estimation tenths of a millimetre) the lengths of 1, 2, . . . 6 ins.

3. On a sheet of ruled quarto manuscript paper measure the distance of each line from the top line.

4. Draw in your note-book a straight line 10 cm. long; divide it into cms. Construct on this a rectangle ABCD, 10 cm.  $\times$  1 cm., and divide it into 10 squares (Fig. 6). Join AC. Then BC will be  $\frac{1}{10}$  of a cm., EG will be  $\frac{5}{10}$  cm., HI will be  $\frac{3}{10}$  cm. Use this scale and a pair of compasses or dividers, to mark points on a line in your note-book,  $\cdot 3$ ,  $\cdot 7$ ,  $2\cdot 3$ ,  $2\cdot 9$ ,  $3\cdot 5$ ,  $4\cdot 1$  cms. apart.

5. Draw a rectangle, 1 inch  $\times$  0.1 inch. Divide it into 10 squares, each 0.1  $\times$  0.1 inch. Draw a diagonal. Use this scale and a pair of compasses to mark points on a line in your note-book,  $\cdot 05$ ,  $\cdot 09$ ,  $\cdot 1$ ,  $\cdot 37$ ,  $\cdot 86$ , and  $3\cdot 42$  inches apart.

6. Find the circumference of (a) a shilling, (b) a halfcrown, by running it along a straight line in your note-book.

7. Draw 2 circles of diameters 5 and 10 cms. Draw also their

diameters. By means of a coin used in No. 6 find the length of each circumference by wheeling the coin round the circumference from a fixed point to the same again.

Divide the length of the circumference by that of the diameter in each case.

8. Measure the length of your working-bench both in English and Metric units. Repeat twice over. Find the mean of your measurements.

9. Measure the length and breadth of the face of a cupboard both in English and Metric units. Repeat twice. Find the mean of your measurements.

Calculate from the mean results stated in Metric units the same lengths stated in English units.

### (ii) Measurement of Area

10. Find the area of a post-card (*a*) in square cm., (*b*) in square mm.

11. Find the area of a visiting-card in square inches and a decimal.

12. Find the area of a filter-paper, and express the answer (*a*) in sq. dm., (*b*) in sq. cm. [Area of a circle =  $\pi r^2$ .]

**The Vernier.**<sup>1</sup>—Fractions of a scale division may be more accurately measured by means of a sliding-scale invented in 1631 by Pierre Vernier, and called a *Vernier*.

**EXERCISE.—To make and use a Scale and Vernier.**

Make a paper scale A B, and divide it into half inches. Mount this on wood (Fig. 7).

Make another paper scale C D, nine half-inches in length ;

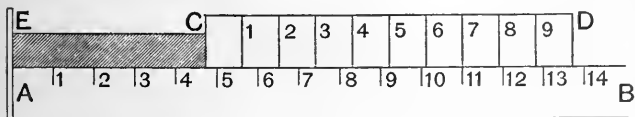


FIG. 7.

divide it into ten equal parts. Mount this on a wooden rod of the same length. C D is the sliding scale or vernier. Now

10 divisions of the vernier = 9 divisions of the scale.

Hence, 1        „        „        „        = 9        „        „

1        „        „        „        falls short of a scale division by 1  
of a scale division.

<sup>1</sup> May be omitted in a first study of the subject.

*To find the length of a Rod E C.*—Place the end E of the rod opposite the 0 of the scale. Its length is 4 + scale divisions. Bring up the end of the vernier against the end C of the rod. In the figure the lines on the vernier do not all correspond with those on the scale. But the seventh line on the vernier is exactly opposite a scale division. Hence the end of the vernier at C must be  $7 \times .1$  scale divisions above the scale division 4.

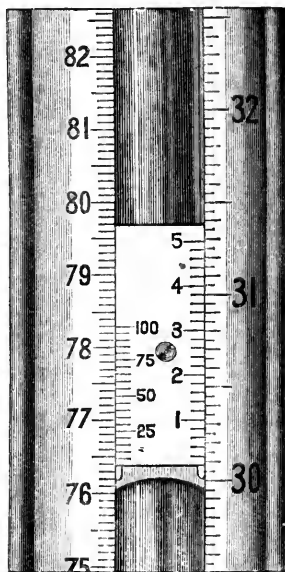


FIG. 8.

Hence the reading is 4.7 scale divisions. But each scale division is .5 inch, and therefore the reading is 2.35 inches.

**The Barometer Vernier.**<sup>1</sup>—There is generally a vernier attached to a barometer. Usually, when the scale is English (as on the right side of Fig. 8), 25 vernier divisions are equal to 24 scale divisions, and 1 scale division is equal to  $\frac{1}{25}$  of an inch.

<sup>1</sup> Omit on a first study of the subject.



If so,

$$\begin{aligned} 1 \text{ vernier division} &= \frac{24}{25} \text{ scale division.} \\ &= \frac{24}{25} \times \frac{1}{10} \text{ or } .048 \text{ inch.} \end{aligned}$$

Hence, 1 vernier division is ( $.050 - .048$ ) or  $.002$  inch less than a scale division.

In order to read the height of the top of the mercury column the vernier must be adjusted so as to be a tangent to the curved mercury surface (*see* p. 64).

What is the relation between a vernier division and a scale division on the left side of Fig. 9, where the scale is French?

#### EXERCISE

Take a piece of zinc, and cut it with scissors so as to be exactly  $\frac{9}{10}$  inch long. Divide it into 10 equal parts, scratching the divisions with a knife. Use this vernier along with a 12-inch rule, divided into tenths, to measure cut slips of cardboard of various lengths.

## CHAPTER IV

### THE MEASUREMENT OF MASS—THE BALANCE

THE mass or quantity of matter in a substance is measured by comparing it with standard masses. These standard masses are usually called "weights." Hence, the measurement of mass is commonly spoken of as "weighing." The measurement of mass is usually carried out with an apparatus called a *balance*.

**EXERCISE.—To make a Simple Balance.**—Fix a clamp firmly to a retort-stand. Take a piece of narrow glass tubing about 40 cm. long. Endeavour to balance it horizontally on the round rod of the clamp. It balances about its mid-point. Tie a thread to each of two 50 gram weights. Suspend each from the glass tube on opposite sides of the clamp by means of a loop on the thread. Find how far each must be placed from the clamp O (Fig. 9) in order that they may balance one another. This only occurs when they are at equal distances from O. If then we suspend two substances at equal distances from O on the two arms, and they

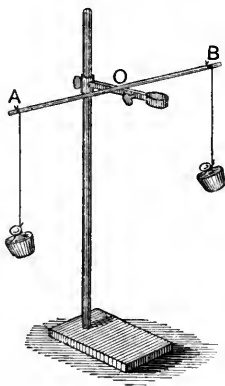


FIG. 9.

balance one another, we may conclude that their masses or "weights" are equal.

Suppose we desire to find the mass or "weight" of a piece of

iron. Suspend it from A. Suspend a small muslin bag from B, such that the length BO = the length AO. Add standard masses or "weights" to the bag until there is equilibrium. Then we may conclude that the weight of the iron is roughly measured by the amount of the standard weights in the bag.

**EXERCISE.**—Balance a 20 gram weight and a 10 gram weight on opposite sides of the clamp. Measure in each case the distance between the point of support and O. What is the relation between these distances?

Does this relation hold :—

*One weight  $\times$  its distance from O = the other weight  $\times$  its distance from O?*

Balance the same weights in fresh positions. Proceed in the same way.

Of all the instruments which are at the present day in the service of the chemist the balance is the most important. To-day the balance is to the chemist as his very right hand, and without it the rapid progress of the last century would have been impossible. The use of the balance has placed chemistry among the exact sciences, and so far as may be chemical problems are now put to the test in the laboratory, and not referred for an answer to the logician in his study. How different our opinion of the balance is from that of Jean Rey, a shrewd physician, who said in 1630 :

"I affirm that the examination of weights which is made by the Balance differs greatly from that which is made by the reason. The latter is only employed by the judicious, the former can be practised by the veriest clown. The latter is always exact ; the former is seldom without deception."

The balance is a delicate instrument, and Jean Rey's clown cannot be allowed to play with it, or it will soon be injured.

**The Student's Balance.**—We shall now describe the balance to be used in almost all the experiments and problems described in this book.

It consists of a beam A B (Fig. 10) which can be raised from or lowered on to a knife edge at C by the handle D. At its two ends are suspended pans E and F. A needle H is attached rigidly at right angles to the beam, and its point is in front of a graduated scale. There are screw nuts at K and G by means of which the lengths of the arms can be slightly altered. The

nuts K and G should be so adjusted that when the pans are empty the needle swings over the same number of degrees to the right and the left of the zero point at the centre of the scale.<sup>1</sup>

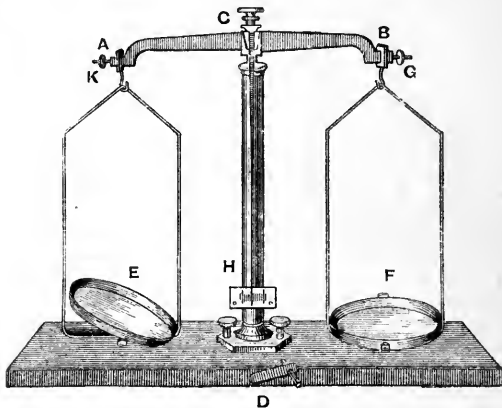


FIG. 10.

Remove the pans from a balance. Examine the means of support, including the knife edges. Make a drawing of the way in which the pan is supported.

**The Weights.**—In a set of weights capable of weighing from '01 to 1000 grams the following individual weights are usually found :—

500	200	100	100
50	20	10	10
5	2	1	1
0'5	0'2	0'01	0'01
0'05	0'02	0'01	0'01

Notice that there are just enough weights to add up to any sum from '01 to 1000 grams.

<sup>1</sup> Several firms sell balances of this type. For laboratory work it is recommended to obtain them fitted with glass cases, whereby the balance is protected from the disturbing influence of air currents, and the still more dangerous effects of corrosive chemical fumes.

**A Rough Balance.**—It is often only desired to know quickly the approximate weights of substances. A very convenient balance is Salter's Letter Balance, No. 1, which will carry 1000 grams. (G. Salter and Co., West Bromwich. 5s. 6d.)

Another convenient balance is one built on what is known as the French type.

For example, if a substance weighing 567·89 grams were weighed the weights would be made up as follows :—

$$\left\{ \begin{array}{l} 500 \\ + 50 + 10 \\ + 5 + 2 \\ + '5 + '2 + '1 \\ + '05 + '02 + '02 \end{array} \right.$$

**The Rider.**—The '01 gram or centigram weight is often a small bent wire called a *rider*, because it can ride astride the balance beam. When placed at different points along the beam its effective weight changes and makes it possible to weigh to milligrams or even to tenths of a milligram. The beam of an accurate balance is divided into ten equal spaces. When the rider is put on the outermost of these it is worth its full weight, '01 gram. On the middle division it acts as '005 gram, and on the innermost division as '001 gram.

**The Platinum Weights,** which are always used in accurate weighing, are differently marked by different makers.

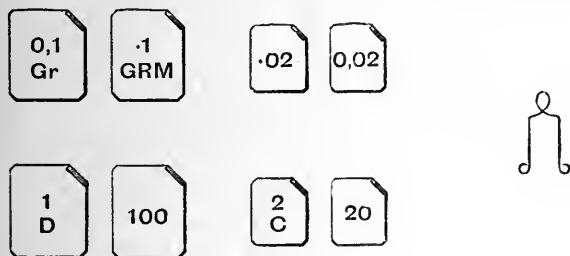


FIG. 11.

The platinum weights shown in Fig. 11 represent :—

'1 gram	'02 gram	'01 gram
= 1 decigram	= 2 centigrams	= 1 centigram
= 100 milligrams	= 20 milligrams	= 10 milligrams.

But if the student knows what to expect he ought to have no difficulty in recognising these and the other platinum weights.

**EXERCISE.**—Remove all the platinum weights from the box.

Arrange them on the squares drawn in Fig. 12. Ask your teacher whether you are right. Rearrange the weights in their box. They should be in order of magnitude. Never force a weight into a hole too small for it. Give each weight a separate space if you can. The bent-up corner should be at the upper right-hand corner. Ask your teacher whether this has been properly done.

·5	·2	·1
·05	·02	·01

FIG. 12.

**The Method of Weighing.**—The following points should be observed :—

1. See that the pointer is swinging equally upon both sides of the scale, when the pans are empty. [Why?]
2. Stop the swinging of the balance whenever the object or the weights are added to or removed from the pans. [Why?]
3. Always place the object to be weighed in the left-hand pan, and the weights in the right-hand pan. [Why?]
4. The position of the observer should be opposite the centre of the scale. [Why?]
5. First add the largest weight necessary, and then try the smaller ones in *descending order of magnitude, omitting none*, but removing those that are too heavy before substituting lighter ones, and arranging the weights on the pan in order of magnitude.
6. When the object is exactly balanced, write down the weight missing from their places in the box, and also the weights on the balance pan, and see that the two agree.

## EXERCISES

## Weighing

1. Find whether the balance swings true, the pointer vibrating to equal distances on each side of the middle point of the scale. Probably it does not. Take a square decimetre of paper. Tear it in two. Letter one half A. Tear the other half in two. Letter one part B. Go on bisecting; letter the weights so made A, B, C, . . . . H. Use these weights to counterpoise the balance until it does swing true.

2. If the error of the balance is less than a centigram use the centigram rider, and without any other weight find what weight would adjust the balance.

3. Pick out weights amounting to 987·65, arranging them neatly on the balance pan, the largest in the centre.

4. Add up the weights on the balance pan.

5. Add up the spaces in the box.

6. Take a set of English *grain* weights, 600 grains to 10 grains, and find the weight of each in grams to the nearest decigram *next below the true weight*. Write answers thus—11·9 +.

7. Repeat exercise 6 to the nearest centigram *next below the true weight*.

8. Cut a piece of lead to weigh exactly 10 grams.

9. Cut a piece of lead to weigh exactly ·1 gram.

10. Weigh newly minted French (or English) silver coins.

[Further exercises in weighing will be found in the succeeding chapters.]

## CHAPTER V

### THE MEASUREMENT OF THE VOLUME OF A LIQUID

**The Units of Volume.**—These are the *litre* for large volumes, and the *cubic centimetre* for small volumes. The litre is equal to 1000 cubic centimetres.

**EXERCISE 1.** Cut out of soap a cube, each edge of which shall measure 1 centimetre. Make a mental note of its size.

It is very important to remember that 1 c.c. of water weighs 1 gram.<sup>1</sup> Consequently, we may regard any given number of grams of water as occupying a volume expressed by the same number of cc. For example, 1 litre or 1000 c.c. of water will weigh 1000 grams.

**EXERCISE 2. To make a Cubic Decimetre.**—Cut out a cardboard cross, 30 cm. by 10 cm. (Fig. 13). Print inscriptions on the four faces as indicated in the figure. Cut the card half-way through along each of the lines A B, B C, C D, D A. Bend back the four squares, and stick their edges together with gummed paper both inside and outside. Let the gum dry. Then warm the hollow cube, and paint it over with melted paraffin.

Counterpoise the cube on a strong balance. Fill it carefully with water. Find the increase in weight. What is the weight of 1 c.c. of the water?

It is now evident that we can readily find the volume of any

<sup>1</sup> This is strictly true only at 4° C. The average temperature of a room is about 15° C. Hence, since water expands slightly between 4° C. and 15° C., 1000 c.c. of water will at 15° C. weigh slightly less than 1000 grams.



vessel by weighing it first empty, and secondly filled with water. For the increase in weight expressed in grams measures its volume expressed in cubic centimetres.

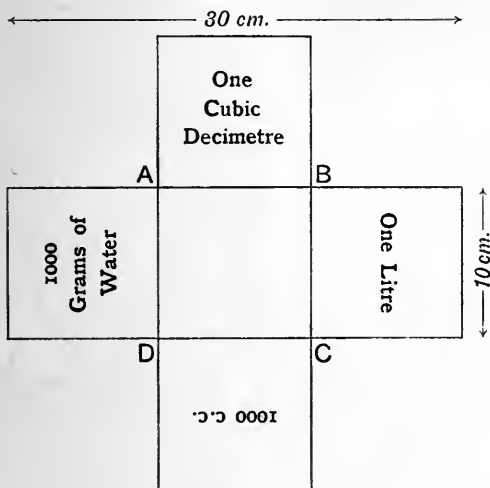


FIG. 13.

## EXERCISES

## Volumes of Liquids

1. Take flasks of 2, 3 and 4 oz. capacity, respectively. With a sharp file make a horizontal mark upon the neck of each. Weigh each flask. Then fill with water up to the mark and weigh again. Express the capacity of each flask in c.c.

Find the capacity of—

2. an ink-bottle.      3. a wine-glass.      4. an egg-cup.

Find what weight of water can be contained in the flasks marked—

5. 100 c.c.      6. 250 c.c.

**Measuring Flasks.**—In carrying out experimental work it is found convenient to have flasks with which to measure out a definite volume of liquid, for instance 100, 250, 500, 1000 c.c.

**EXERCISE 3. To mark a Flask to contain 100 c.c.—** Take a dry 3-ounce flask with a narrow neck. Weigh empty. Then add water till there is an increase in weight of 100 grams. Allow the flask to stand until any water on the side of the neck

has drained down. With a sharp file moistened with water scratch a horizontal line on the neck to mark the lowest point of the curved surface bounding the water and the air.

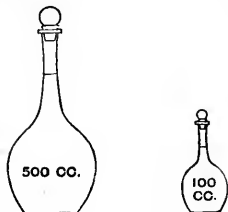


FIG. 14.

The measuring flasks most often used are litre,  $\frac{1}{2}$ -litre,  $\frac{1}{4}$ -litre, and 100 c.c. flasks (Fig. 14). They are usually fitted with ground glass stoppers, and when filled up to the circular line etched upon their necks

contain at 15° C. the volumes which are marked upon them.

**Measuring Jars.**—These are tall cylindrical vessels, graduated in cubic centimetres, and holding 100, 200, 250, 500 or 1000 c.c. (Fig. 14A).

**Pipettes.**—Flasks and measuring jars measure the liquids they *contain*, but it is extremely useful to have vessels which will measure out, or *deliver*, accurately a small definite volume of liquid.

A pipette is a glass tube open at both ends, upon which a cylindrical bulb has been blown (Fig. 15). There is a circular mark on the stem. When filled with liquid up to the mark the pipette will deliver exactly that volume which is marked upon it.

**EXERCISE 4. To use a Pipette.**—Place the pointed end of the pipette in water. Suck up the water till above the mark. Quickly close the upper end with the moist forefinger. Pressing gently, rotate the pipette so that the water very slowly flows out until the water-level reaches the mark. Then press tightly. The flow ceases. Transfer the pipette to a beaker, remove the finger, and let the liquid flow out. As soon as the pipette appears to be empty, let it drain for five seconds, with the tip just below the surface of the water in the beaker. The drop of water which is still in the pipette is *left*, and not blown out.



FIG. 14A.

**Burettes.**—A burette is a glass tube with a glass tap at the lower end, used to deliver any desired volume of liquid (Fig. 16). The tube is graduated in cubic centimetres and tenths of a cubic centimetre.

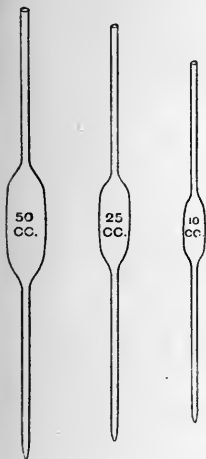


FIG. 15.

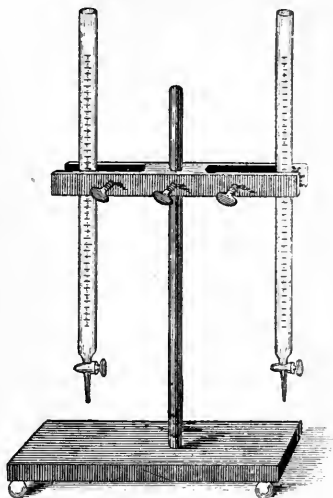


FIG. 16.

**EXERCISE 5. To prepare a Burette for use.**—Fix the burette in a clamp. Remove the tap, and grease it with a very little vaseline. Replace it, and twist a rubber band round it, so as to prevent it falling out of place. Close the tap. Pour water into the tube through a small funnel until its level is above the top of the scale. Place a beaker underneath the tube. Turn the tap, and let the water run out, so that no air is left in the burette below the tap. The burette is now ready to deliver a measured volume of water.

If the figures and scale-divisions on a burette are not easy to read, they may sometimes be made more distinct by rubbing them over with a piece of waxy blue crayon.

**Burette-Reading.**—The surface of a liquid in a burette is curved, and the curve is called a *meniscus*; it is always the *lowest*

*point* of the meniscus which is measured. Note the following points :—

(1) The eye must be on the same level as the curve, otherwise the curve is seen displaced above or below the true position, an effect which astronomers call *parallax*. The level position is best obtained by looking straight past the burette at some shelf known to be on the eye level, or by looking past the curve at its image in a looking-glass.

(2) The extreme edge, or *lowest position*, of the meniscus is, however, sometimes hard to see, but it can be made clear. On looking into an aquarium from below the surface-line goldfish are seen reflected above the surface. This phenomenon of reflection from the inside of a surface is called *internal reflection*, and it is this which may be made use of to render the meniscus distinct. If the background be bright, the meniscus is made dark by holding a dark object *just behind and just below the surface*. The blackened edge of a card slipped over the burette tube answers admirably. At night, however, when the background is dark, a gas flame may be put on the bench below to illuminate the meniscus.

(3) The graduations are decimal, each fifth and tenth line being longer than the rest. Notice the figures opposite the divisions. Find whether they count upwards or downwards. Find how many c.c. each scale-division stands for.

## EXAMPLES VII

### Burette-Reading

1. Read the position of the liquid in the figure of a tube shown below (Fig. 17), taking the two long lines W and X to represent respectively :—

A	B	C	D	E	F	G	H
30	40	7	800	120	130	250	250
and 20	50	6	900	100	150	200	300

For instance, the two longest lines represent 30 c.c. and 20 c.c. respectively in question A, 40 c.c. and 50 c.c. in question B, and so on.

2. Sometimes the lowest point of the meniscus is not exactly opposite

to any of the scale-divisions; tenths or other fractions of a division should then be *estimated* (see page 17).

Read the position of the liquid shown in Figure 18, estimating the



FIG. 17.

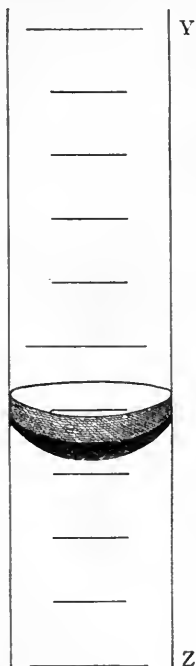


FIG. 18.

position of the curve in tenths of a scale division, taking the two long lines Y and Z to represent respectively:—

J	K	L	M	O	P	Q	R
1	14	10	30	40	700	800	0
and 2	13	11	20	50	600	900	1000

## EXERCISES

## Pipettes and Burettes

- (1) Suck up water in a pipette. Set to the mark.
- (2) Pour water into a burette. Set it exactly at 10 c.c.
- (3) Pour some water out of the burette. Read the new position of the liquid.
- (4) Weigh an empty beaker. Measure into it 10 quantities of water of 10 c.c. each by means of a pipette. Weigh again. What is the gain in weight?
- (5) Choose a 100 *grain* pipette and a 50 c.c. burette, graduated to tenths of 1 c.c. Set the liquid in a burette to 50. Measure 100 grains of water in the pipette, and run it into the burette. Read the level now. Run in another 100 grains, read again, and so on until you read the top of the burette. Take the difference between each two consecutive readings. These differences ought to be equal, for each should be the number of c.c. occupied by 100 grains of water. If the differences are not exactly equal, find out where you have made a mistake. Repeat the exercise with still greater care, estimating tenths of a division.

State your results thus :—

## EXPT. 1. Measuring to 1 c.c.

## Differences.

0.	50	}	. . 6.3
1.	43.7		
		}	. . 6.5
2.	37.2		
		}	. . 6.4
3.	30.8		
		}	. . 6.5
4.	24.3		
		}	. . 6.5
5.	17.8		
		}	. . 6.4
6.	11.4		
7.	. . . . .		

## EXPT. 2. Estimating .01 c.c.

## Differences.

50	}	. . 6.49
43.51		
	}	. . 6.46
37.05		
	}	. . 6.49
30.56		
	}	. . 6.46
24.10		
	}	. . 6.48
17.62		
	}	. . 6.48
11.14		
	}	. . 6.45
4.69		

(6) Pour water into a burette. Set it exactly at 0 c.c. Weigh an empty beaker. Run into it exactly 5 c.c. Weigh. Run in another 5 c.c., weigh again, and so on until you have run out 50 c.c. Take the difference between each two consecutive weighings. These differences should be equal.

(7) You are provided with a measuring jar, or with a beaker and a burette and some marbles. Devise a method of finding their volume.

In Exercises 1-3 *ask your teacher whether you are right or not.* In Exercises 4-6 you ought to be able to detect your own errors without help.

## CHAPTER VI

### THE MEASUREMENT OF TEMPERATURE

**Heat and Temperature.**—Every one can attach some meaning to the word *Heat*. The sensation of heat is familiar to all. And yet we shall often make mistakes in judging whether a substance is hot or cold by our sensations. For the same room may at the same time appear hot to one man and cold to another. Indeed, the two hands may judge differently of the same substance; test this by placing one hand for a few seconds in hot water and the other in cold, and then place both together in lukewarm water: one hand feels cold and the other warm. It is clear then that the sensations of heat and cold in our bodies do not necessarily tell us anything of the actual state of a substance with regard to heat or cold.

Nevertheless there is no room for doubt that different states of a body with regard to heat and cold can exist, and we shall proceed to consider how these different states or conditions can be measured. This we can do without stopping to find out exactly what heat itself is.

**Temperature.**—If a poker be placed in a fire the end outside the fire quickly becomes warm, and we understand that heat has passed from the hot coals into the poker. If the red-hot poker be placed in a bucket of cold water the poker will be rapidly cooled, and we say that heat has passed from the poker into the water. Many instances are met with every day of heat passing from one body to another; such bodies are said to be at different *temperatures*, and the heat passes from the body



which is said to have the higher temperature to that which is said to have the lower temperature.

A difference in temperature between two bodies may be compared to a difference in level between the water in two cisterns. When two bodies are so placed that heat can flow between them, the flow of heat will always take place from the body which has the higher temperature to that which has the lower temperature, just as water will always flow from a higher level to a lower. Indeed, temperature may be defined as *that condition of a body upon which depends the flow of heat to or from other bodies*. It is important also to observe that the passage of heat from a hot to a cold body will continue until they reach a common temperature, just as water is said to find its own level.

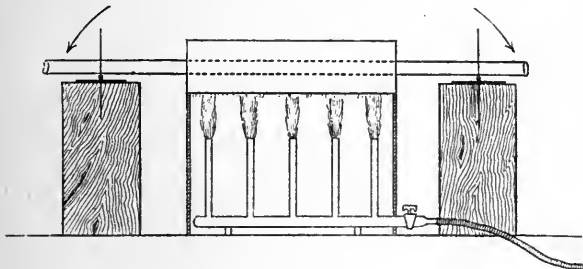


FIG. 19.

**Preliminary Idea of a Thermometer.**—When heat is applied to any body it may produce many effects, but none is more noticeable than change in volume. In by far the greater number of cases a body expands or increases in volume when it is heated.

**EXPT. I. To show that Iron expands when heated.**<sup>1</sup>—Support an iron gas pipe at its two ends upon two needles, resting upon horizontal glass plates, and let the central portion of the pipe lie within (but not touching) a gas furnace (Fig. 19). Fix to each needle a light straw to serve as an index. Now light the furnace. In a short time each needle begins to roll outwards turning the straw. This movement must be caused by the lengthening (expansion) of the bar. Note that the hotter the

<sup>1</sup> Lecture Table Experiment.

pipe becomes the greater is the expansion, and the consequent movement of the straws. Allow it to cool. Note how the needles roll back.

**EXPT. 2. To show that Water expands when heated.**—Take a glass flask. Select a rubber stopper with one hole which fits it closely. Pass a narrow glass tube, 50 cm. in length, through the cork so that the end of the tube is flush with the lower surface of the stopper (Fig. 20). Fill the flask quite full with water coloured with indigo. Push the stopper with the tube into the neck of the flask so that all air is expelled, and the liquid rises a short way up the tube. Mark

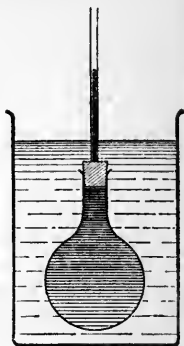


FIG. 20.

the level of the liquid with a piece of gummed label. Immerse the whole of the flask in a vessel containing hot water. What do you observe? The liquid in the tube first falls slightly; then it rises steadily. Why? The fall at first may have been due to a contraction of the water, or to an expansion of the glass flask. The heat had hardly time to reach the water, so probably the fall was due to the flask expanding. The second change—the slow steady rise—is due to the gradual heating and expansion of the water within the flask.

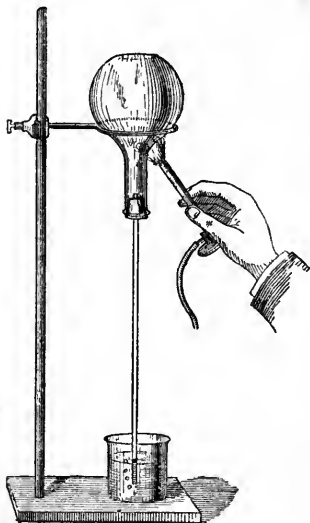


FIG. 21.

**EXPT. 3. To show that Air expands when heated.**—Close a glass flask with a rubber stopper through which is passed a long glass tube. Invert the flask and support it so that the open end of the tube dips under coloured water con-

tained in a beaker (Fig. 21). Warm the flask first with the hand. A few bubbles escape. Then warm the flask with a flame. The enclosed air bubbles out much faster. Remove the source of heat. Water rises up the tube, showing that the air contracts on cooling.

In these three experiments we have illustrations of the expansion by heat of a solid, of a liquid, and of a gas. In each case the expansion indicates the flow of heat into them from sources which are at a higher temperature. The simple instrument described in Experiment 2 may be used to compare the temperature on different days, or in different parts of a house, provided in each case the instrument is left sufficiently long for the passage of heat to continue until the instrument and the room have acquired a common temperature. If the glass tube be graduated, that is to say divided into equal divisions and numbered, temperatures may be stated in terms of this scale. Such an apparatus is a simple instrument for recording and measuring temperatures—in other words it is a rough *thermometer*. It may be remarked that Galileo in the year 1597 was probably the first to employ a thermometer.

### The Mercurial Thermometer.

The ordinary mercurial thermometer is essentially just such an instrument as was used in Experiment 2. Mercury, however, is used as the liquid instead of water. The principal advantages of mercury over water are that it does not *freeze* until a very low temperature, and secondly that it does not *boil* until a very high temperature. The mercurial thermometer consists of a very fine thick-walled glass tube, which ends in a round or cylindrical bulb (Fig. 22).

#### The Fixed Points on a Thermometer.—

In order that a thermometer may be of practical use its stem must be graduated to allow of comparative readings being taken.

In graduating the thermometer it is customary in the first place to mark upon its stem the points at which the mercury stands when it is heated to two standard temperatures. It was Newton who suggested that for these *fixed points*, as they are called, the



FIG. 22.

melting temperature of ice, which is found to be very constant, and the temperature of steam arising from boiling water, which also is constant so long as the pressure of the air remains the same, should be employed.

**EXPT. 4. To Mark the Boiling Point on a Thermometer.**—To a flask A containing some water attach by means of a cork a glass tube B, and surround the upper portion of this by a wider tube C, into which is fused a small side tube at D (Fig. 23). Fix the thermometer within the inner tube by means of the cork E, and boil the water vigorously. The thermometer is now surrounded by steam. Observe carefully the point to which the mercury rises. This point is marked one hundred degrees or  $100^{\circ}$  on a Centigrade and  $212^{\circ}$  on a Fahrenheit thermometer.

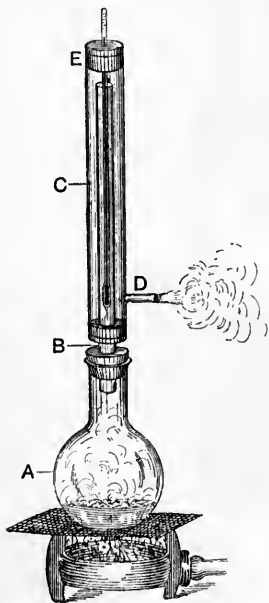


FIG. 23.

**EXPT. 5. To Mark the Freezing Point on a Thermometer.**—Fill a funnel with pounded ice. Support it on a tripod. Plunge the thermometer into the mass (Fig. 24). Observe the point at which the mercury becomes stationary. This is marked  $0^{\circ}$  on a Centigrade and  $32^{\circ}$  on a Fahrenheit thermometer.

**Scales of Temperature.**—

The difference in temperature between these two fixed points is large. It is customary to subdivide the interval between them according to one or other of 3 scales, named the Fahrenheit, Centigrade, and Réaumur scales; we need only consider the first two.

On the Fahrenheit scale the freezing point is marked  $32^{\circ}$  and the boiling point  $212^{\circ}$ , while on the Centigrade scale they are marked  $0^{\circ}$  and  $100^{\circ}$  respectively (Fig. 25). The interval between these points is therefore divided into 180 equal divisions

or degrees on the Fahrenheit and into  $100^{\circ}$  upon the Centigrade scale.

Hence,  $180$  Fahrenheit degrees =  $100$  Centigrade degrees.

or	9	“	“	=	5	“	“
Therefore,	1	“	“	=	$\frac{5}{9}$	“	“
and	$\frac{9}{5}$	“	“	=	1	“	“

On the Fahrenheit scale zero is  $32^{\circ}$  below the freezing point.

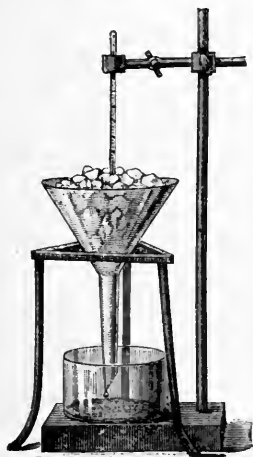


FIG. 24.

### Conversion of Thermometric Scales.—

(1) Find the Centigrade reading corresponding to  $80^{\circ}$  F.

$80^{\circ}$  F. is  $(80^{\circ} - 32^{\circ})$  or  $48^{\circ}$  F. above the freezing point.

$48^{\circ}$  F. is  $48^{\circ} \times \frac{5}{9}$  C. or  $26^{\circ} \cdot 6$  C. above the freezing point, which is  $0^{\circ}$  on the C. scale.

Hence, the required reading is  $26^{\circ} \cdot 6$  C.

(2) Find the Centigrade reading corresponding to  $20^{\circ}$  F.

$20^{\circ}$  F. is  $(32^{\circ} - 20^{\circ})$  or  $12^{\circ}$  F. below the freezing point.

$12^{\circ}$  F. is  $12^{\circ} \times \frac{5}{9}$  C. or  $6^{\circ} \cdot 6$  C. below the freezing point, which is  $0^{\circ}$  on the C scale.

Hence, the required reading is  $-6^{\circ} \cdot 6$  C.

(3) Find the Fahrenheit reading corresponding to  $15^{\circ}\text{C}$ .

$15^{\circ}\text{C}$ . is  $12^{\circ} \times \frac{9}{5}$  F. or  $27^{\circ}\text{F}$ . above the freezing point, which is  $32^{\circ}$  on the Fahrenheit scale.

Hence the required reading is  $(27^{\circ} + 32^{\circ})\text{F}$  or  $59^{\circ}\text{F}$ .

**Mechanical Interpolation.**—The conversion of thermometer scales is only one instance of the mathematical process

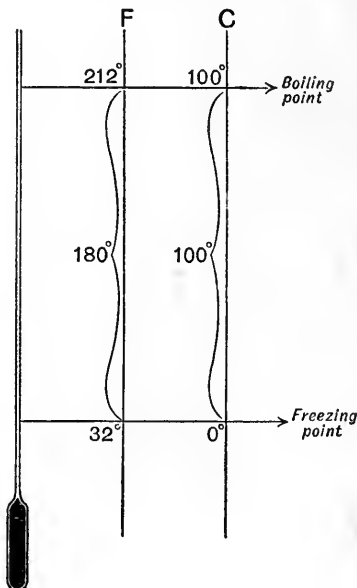
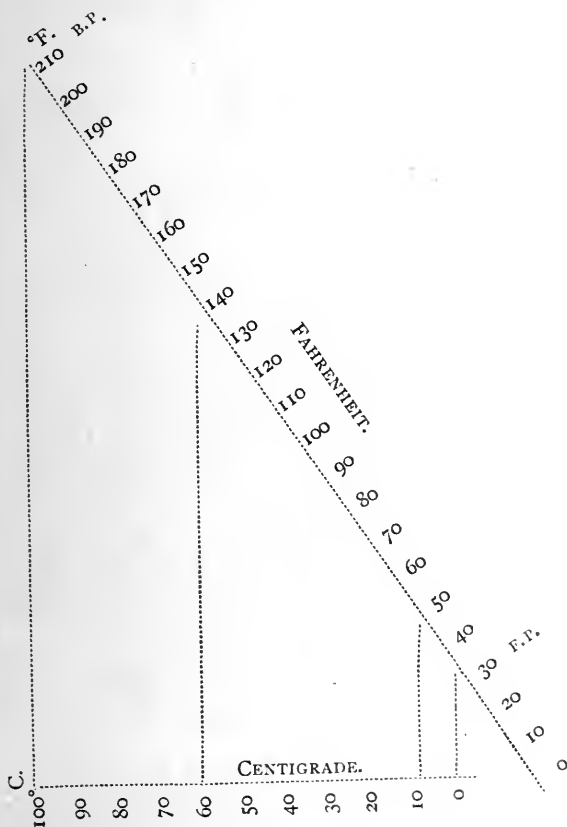


FIG. 25.

called *interpolation*. Other instances occur in making out railway time-tables, in arranging prices and sizes of goods, and in the use and preparation of logarithmic, astronomical, and other tables.

Such processes of interpolation, for instance, that of temperature conversions, may be carried out mechanically. Thus, take a sheet of manuscript paper ruled with equidistant lines. Along the edge number the consecutive lines 0, 10, 20, . . . 100

for Centigrade temperatures. On a narrow strip of the same paper write 30, 40, . . . 210 for the Fahrenheit temperatures. Lay the strip across the sheet so that  $212^{\circ}$  F. falls on the line marked  $100^{\circ}$  C., and  $32^{\circ}$  F. on the line  $0^{\circ}$  C. Any temperature on the Fahrenheit scale will now be found exactly opposite the corresponding Centigrade temperature; *e.g.*,  $10^{\circ}$  C. =  $50^{\circ}$  F. and  $140^{\circ}$  F. =  $60^{\circ}$  C.



**To Read the Thermometer.**—Although the interval between the fixed points on a Centigrade thermometer is divided into only 100 degrees, it is nevertheless possible by means of the eye to mentally subdivide each degree into tenths, and therefore to read any given temperature to the nearest *tenth* of a degree.

Fig. 26 shows a piece of a thermometer stem. The graduations stand for single degrees. The end of the mercurial column lies between  $11^{\circ}$  and  $12^{\circ}$ . Subdividing the eleventh degree by the eye, the reading is  $11.4$ .



FIG. 26.

## EXAMPLES VIII

**Interpolation—Conversion of Therometric Scales****(i) INTERPOLATION.**

1. A barometer has two scales attached to it both counting from the same zero. 760 on one scale is opposite 30 on the other. What number on the second scale is opposite 750 on the first?

2. In a certain street the lamp-posts and the houses do not count from the same point. The third lamp-post is opposite No. 21, and the seventh lamp-post is opposite No. 45. What number is opposite the tenth lamp-post?

3. On a certain railway journey by express K, L, and M are passed at 12.3, 12.21, and 12.27. A slower train passes K and M at 1.4 and 1.40. At what time does it pass L?

4. In a certain examination A got 150 marks, Q 70, and Z 30. The marks were recalculated so as to give A 100 and Z 0. How many did Q get?

5. A football professional is paid partly by a fixed monthly salary and partly by a bonus on every match won. He got £11 10s. when 7 matches were won in a month, and £14 for 12 matches won in a month. How much will he get for winning 5 matches in a fortnight?

6. If the sun sets at 6.10 P.M. on September 16, and at 5.54 on September 23, find the time of sunset on September 21.

7. A solution containing 40 per cent. sulphuric acid has the density 1.306, and a 50 per cent. solution has the density 1.398. What is the density of a 46 per cent. solution?

**(ii) CONVERSION OF THERMOMETRIC SCALES.**

What Fahrenheit temperatures are equal to

8.  $15^{\circ}\text{C.}$

9.  $5^{\circ}\text{C.}$

10.  $80^{\circ}\text{C.}$



What Centigrade temperatures are equal to

11.  $50^{\circ}$  F.

12.  $95^{\circ}$  F.

13.  $122^{\circ}$  F.?

Express on the Fahrenheit scale—

14.  $20^{\circ}$  C.

15.  $-5^{\circ}$  C.

16.  $4.5^{\circ}$  C.

Express on the Centigrade scale—

17.  $40^{\circ}$  F.

18.  $14^{\circ}$  F.

19.  $-4^{\circ}$  F.

20. On the Réaumur scale the freezing point is  $0^{\circ}$  and the boiling point  $80^{\circ}$ . What Réaumur temperature is equal to  $59^{\circ}$  F.?

21. What Fahrenheit temperature is equal to  $16^{\circ}$  R.?

## EXERCISES

### Thermometry

1. Place a Fahrenheit and a Centigrade thermometer in a beaker of cold water. Take the C. reading, and convert it into F. degrees. Does the number agree with the F. thermometer?

2. Warm the water in the beaker, and when the C. thermometer reads about  $40^{\circ}$ , read both thermometers accurately, and as nearly as possible at the same time. Convert the C. reading to the F. scale. Does the result agree with the F. reading?

3. Heat the water longer, until the C. thermometer reads about  $75^{\circ}$ , and then read both thermometers. Convert the F. reading to the C. scale. Does the result agree with the C. reading?

4. Now boil the water in the beaker, and read the C. thermometer. Arrange two columns in your note-book headed **Time** and **Temperature**, thus—

TIME.			TEMPERATURE.
3 hr.	20 m.	0 sec.	$100^{\circ}.2$ C.
3	20	30	
3	21	0	
3	21	30	

Put down the temperature of the boiling water opposite to the first time reading. Then look at the second hand of your watch, or at the laboratory clock, and when it is at the minute extinguish the flame. Then at intervals of half-minutes take the readings of the C. thermometer to the first decimal place. Take 15 readings and plot a time-temperature curve of the rate of cooling on squared paper.

5. Half fill a flask with water. Add four or five grams of common salt. Heat the flask until the water boils. Hold the bulb of a C. thermometer in the steam *above* the boiling liquid. What is the highest temperature it records?

6. Place the bulb of the thermometer *in* the boiling liquid. Is the reading the same?

7. Pound some ice in a mortar. Add a small handful of common salt. Mix thoroughly. Read the lowest temperature reached.

## CHAPTER VII

### THE MEASUREMENT OF THE RELATIVE DENSITIES OF SOLIDS AND LIQUIDS

#### (A.) THE RELATIVE DENSITIES OF LIQUIDS

No one can fail to be aware that there are many different kinds of liquids. At the mention of mercury, water, oil, or methylated spirits, we remember also at once that they are not all of the same weight. How can we compare the weight of one liquid with that of another? Obviously by weighing in turn one and the same volume of each liquid.

**EXPT. 1. To Find the Weight of 100 c.c. of Water at the Temperature of the Room.**—Take a small dry beaker.<sup>1</sup> Weigh. Measure into it by means of a pipette 50 c.c. of water. Weigh. Thus:—

Weight of beaker	...	...	...	...	15.4 grams.
Weight of beaker + 50 c.c. water	...	...	...	...	65.3 „
Then, Weight of 50 c.c. water	...	...	...	...	49.9 „
Hence, Weight of 100 c.c. water...	...	...	...	...	99.8 „

#### EXERCISES

Find the weight of 100 c.c. of (1) strong hydrochloric acid; (2) oil of turpentine; (3) methylated spirits; (4) strong sulphuric acid; (5) milk; (6) glycerine; (7) sea-water.

---

<sup>1</sup> A 100 cc. measuring flask, if at hand, is more suitable.

Compare your results with the following numbers which were obtained in the same way :—

Methylated spirits	...	...	...	...	81.6	grams.
Oil of turpentine	...	...	...	...	86.3	„
Sea-water	...	...	...	...	102.5	„
Milk	...	...	...	...	102.9	„
Strong hydrochloric acid	...	...	...	...	115.5	„
Glycerine	...	...	...	...	126	„
Strong sulphuric acid	...	...	...	...	185.4	„

It is evident then that equal volumes of these different liquids contain different weights of matter. This is usually expressed by saying that some liquids are *denser* than others, or that they possess different *densities*.<sup>1</sup>

It has been found convenient to choose water as a standard liquid, and to compare the densities of other liquids with that of water. We can readily obtain from the results of our experiments numbers which express the *density* of any liquid *relative* to that of water.

**Definition of Relative Density.**<sup>2</sup>—*The relative density of a substance is a number which expresses the weight of a given volume of the substance divided by the weight of the same volume of water.*

For instance :—

$$\begin{aligned}
 \text{Relative density of milk} &= \frac{\text{weight of 100 c.c. of milk}}{\text{weight of 100 c.c. of water}} \\
 &= \frac{102.9}{99.7} \\
 &= 1.032
 \end{aligned}$$

In this way we find the following relative densities, or specific gravities, as they are often termed :—

Methylated spirits	...	...	...	...	...	0.818
Oil of turpentine	...	...	...	...	...	0.865
Water	...	...	...	...	...	1.000

<sup>1</sup> **Definition of Density.**—*The density of a substance is defined as the mass of unit volume.*

Hence, for example, from the results above,

$$\begin{aligned}
 \text{density of methylated spirits} &= 0.816 \\
 \text{density of strong sulphuric acid} &= 1.854
 \end{aligned}$$

<sup>2</sup> Often called Specific Gravity.

Sea-water ... ..	1.027
Milk ... ..	1.032
Hydrochloric acid ... ..	1.159
Glycerine ... ..	1.263
Sulphuric acid ... ..	1.858

It is important to observe that determinations with different samples of any one liquid, for example oil of turpentine, give very nearly the same numbers. This fact leads us to suspect that any definite liquid has a relative density which is a constant property of it (at a given temperature). So long as the liquid is not mixed with any other liquid its relative density appears to remain constant, and to have a fixed value. Consequently, it is possible to distinguish between liquids which may otherwise resemble one another by a reference to their relative densities. For instance, the adulteration of milk with water may be detected by finding the lowering of relative density which results from the admixture.

#### **Influence of Temperature on Relative Density.—**

We know that water and most other liquids expand when they are heated. Consequently, a given volume of hot water ought to weigh less than the same volume of cold water.

**EXPT. 2. To show that Hot Water weighs less than Cold Water.**—Heat some water till it boils. Weigh a small flask with a cork. Remove the flame, and rinse out a 20 c.c. pipette with the hot water. Take the temperature of the water, and then measure 20 c.c. of the water into the flask and close it with the cork. Weigh again, and calculate the weight of 100 c.c. of the hot water. Repeat at lower temperatures, and tabulate the results obtained.

Do given volumes of other liquids also weigh less when hot than when cold? This can readily be tested in the case of glycerine, for example, and it will probably be found to be so.

How will this effect of heat affect the relative density of a liquid? If when expressing the relative density of a liquid at a given temperature we divide the weight of a given volume of the liquid by the weight of the same volume of water at some *fixed* temperature, then the relative density of the liquid will become less and less as the temperature of the liquid rises.

It is better, however, to divide the weight of a given volume of the liquid by the weight of the same volume of water at the

*same* temperature as the liquid. In the case of both the liquid and the water expanding when heated at the same rate, the number expressing the relative density of the liquid will remain constant. But in general this will not be so, and the number expressing the relative density will alter slightly with change of temperature.

### Relative Densities of some Common Liquids.—

Ether ... ..	73
Alcohol ... ..	80
Petroleum ... ..	8+
Ammonia ... ..	88
Olive oil ... ..	9
Water ... ..	1'000
Milk ... ..	1'03
Sea-water ... ..	1'026
Glycerine ... ..	1'26
Sulphuric acid ... ..	1'85
Mercury ... ..	13'5+

### EXERCISES

#### Relative Densities of Liquids

1. Measure into a weighed beaker from a burette or pipette 10 c.c. of water. Weigh. Run in 10 c.c. more. Weigh. Run in another 10 c.c. Weigh. Run in another 10 c.c. Weigh. Add another 10 c.c. Weigh.

2. Repeat No. 1 with methylated spirits.

3. Calculate the relative density of methylated spirits from the numbers obtained for the weights of

- (1) 10 c.c. spirit and 10 c.c. water.
- (2) 20 „ „ 20 „ „
- (3) 30 „ „ 30 „ „
- (4) 40 „ „ 40 „ „
- (5) 50 „ „ 50 „ „

4. Repeat Ex. 1 with sea-water, and calculate its relative density.

5. Introduce 10 grams of common salt into a flask. Add about 100 c.c. of water, and shake until the salt has dissolved. Pour some of the solution into a burette. Proceed as in Ex. 1. Find the average weight of 1 c.c. salt solution.

6. Add 20 c.c. of water to 20 c.c. of methylated spirit. Measure the volume of the mixture. Calculate the weight of the mixture from Ex. 1 and 2. Calculate also the average weight of 1 c.c. of the mixture.

7. Measure 10 c.c. of the mixture made in Ex. 6 into a weighed beaker. Weigh. Find the weight of 1 c.c. Does your result agree with the weight calculated in Ex. 6?

8. Calculate the relative density of the mixture of methylated spirits and water.

9. Heat some water to about  $90^{\circ}\text{C}$ . Weigh a flask with a cork. Read the temperature of the water. Rinse out a 20 c.c. pipette with the hot water. Measure 20 c.c. into the flask, and close it with the cork. Weigh. Calculate the weight of 100 c.c.

10. Repeat Ex. 9, using ice-cold water.

## EXAMPLES IX

### Relative Densities of Liquids

1. Two beakers are placed one on each pan of a balance, and shot are added until the beam swings evenly. Then 5 c.c. of mercury are placed in one beaker, and it is found that 68.8 c.c. of water have to be added to the other beaker to restore the balance. Calculate the density of mercury relative to water, and the density of water relative to mercury.

2. A pipette contains 30.86 grams of spirits of relative density 0.81. What is the volume of the pipette?

3. What weight of sulphuric acid of relative density 1.86 will the same pipette contain?

4. 20 c.c. of a liquid A of relative density 1 are added to 20 c.c. of a liquid B of relative density 0.81. What is the weight of the mixture?

5. The mixture referred to in Ex. 4 measures only 38.5 c.c., owing to contraction having taken place. What is the relative density of the mixture?

6. A flask weighs 25 grams when empty, 125 grams when full of distilled water, and 127.6 when full of sea-water. Find the relative density of the sea-water.

7. How many c.c. of sulphuric acid (1.85) must be measured out to obtain 111 grams of it?

8. Alcohol (.8) and sulphuric acid (1.8) are to be mixed in the proportion 1 : 3 by weight. What volume of alcohol should be taken with 200 c.c. of acid?

9. A salt solution has the relative density 1.025; how many c.c. of water must be added to 1 litre of it to reduce its relative density to 1.020?

10. Find the relative density of a mixture of 100 c.c. water with 100 c.c. spirits of relative density 0.81, assuming that the liquids contract by 2 per cent. of their original volumes on mixing.

11. A 200 c.c. flask weighs 300 grams when full of water. What will it weigh when full of mercury (13.5)?

12. A bottle weighs 20 grams. Full of water at 4°C. it weighs 70 grams. When filled with water at 65°C. it only weighs 69 grams. Find the relative density of the hot water.

13. A tube 30 cm. long weighs 15 grams when empty, and 96.6 grams full of mercury (13.6). Find in sq. mm. the area of the cross section of the tube.

---

## (B.) THE RELATIVE DENSITIES OF SOLIDS

All that has been said of liquids as regards the differences in density which are found to exist, is of course true also of solids. Although it would appear at first sight more difficult to measure the relative densities of solids, there are several ways in which it can be carried out. Remember that it is required to compare the weight of a solid with the weight of an equal volume of water. How can we find the weight of an equal volume of water? We can readily measure the volume of some regular solids such as a cube, a cylinder, or a bar, but how can we find the volume of a number of leaden shot, a heap of sand, or a bundle of steel screws? We shall describe three methods which may be followed.

### (i) First Method:—

**EXPT. 2. To find the Relative Density of a Glass Stopper.**—Weigh a solid glass stopper, say 12.5 grams. How can its volume be found? The shape is not regular enough to be found by mensuration. But we can find how much water the stopper can displace.

Choose a glass measuring jar, the narrowest into which the stopper will easily slide. Partly fill the jar with water. Say it stands at 55 c.c. Carefully drop in the stopper. The level rises



to, say, 60 c.c. Then the volume of the stopper is (60 - 55) or 5 c.c. Then—

$$\begin{aligned}\text{Relative density of the glass} &= \frac{\text{weight of the glass}}{\text{weight of an equal volume of water}} \\ &= \frac{12.5 \text{ grams}}{5 \text{ grams}} \\ &= 2.5\end{aligned}$$

(ii) **Second Method.**—This will be made clear by an example of its application.

**EXPT. 3. To find the Relative Density of Lead Shot.**  
—Weigh a small dry measuring flask with a mark on the neck. Introduce 40 to 50 grams of leaden shot. Weigh again. Cover the shot with water. Shake well to dislodge any air bubbles. Add more water until the lowest point of the meniscus is exactly on a level with the mark. Weigh again. Empty out the shot and weigh the flask filled up to the line with water alone. We have now all the necessary data. Thus in an experiment :—<sup>1</sup>

Weight of empty flask	...	...	...	...	...	35 grams.
Weight of flask and lead	...	...	...	...	...	79 „
Weight of flask and water	...	...	...	...	...	135 „
Weight of flask and lead, filled up with water	...	...	...	...	...	175 „

Hence,

Weight of water alone filling the flask	...	...	...	...	...	= 135 - 35
						= 100 grams.
Weight of lead alone	...	...	...	...	...	= 79 - 35
						= 44 grams.

Also,

Weight of water filling up the flask when the						
lead is present	...	...	...	...	...	= 175 - 79
						= 96 grams.

Hence,

Weight of water displaced by the lead	...	...	...	...	...	= 100 - 96
						= 4 grams.

$$\begin{aligned}\text{But, relative density of lead} &= \frac{\text{weight of lead}}{\text{weight of an equal volume of water}} \\ &= \frac{44}{4} \\ &= 11\end{aligned}$$

<sup>1</sup> All decimals are omitted for the sake of clearness in the example quoted. In actual practice account should be taken of them.

**Principle of Archimedes.**—There is another method of determining the relative densities of solids, which depends upon a fact discovered about 220 B.C. by the philosopher Archimedes of Syracuse. It is said that Hiero, King of Syracuse, sent a lump of gold to a goldsmith's to be made into a crown. When the crown was delivered he suspected that, though of the right weight, it contained nevertheless an undue proportion of silver. So Archimedes was asked to find out the truth of the matter. Turning the matter over in his mind it is said that he went to have a bath, and noticing the rise of the water he jumped out in great excitement and ran naked through the streets of the city crying "Eureka, Eureka, I have discovered it!" But suppose we make the following experiment.

**EXPT. 4. To try whether a Brass Cylinder weighs the same in Water as in Air.**—Take a brass cylinder. Suspend it by means of a silk thread under one pan A of a balance<sup>1</sup> of the type indicated in Fig. 27, and inside a narrow glass measuring jar,<sup>2</sup> so that the cylinder can rise and fall without touching the jar. Place a small beaker on the pan A, and counterpoise the whole by weights added to the pan B. Remove the jar, and unhook the cylinder. Pour about 50 c.c. of water into the jar, and read its level. Suspend the cylinder by the thread in the water. The level of the water is of course raised. Read the new level. The difference between the two readings gives the volume of the cylinder. Take the cylinder out of the water, and place the jar underneath the arm of the balance. Again suspend the cylinder from the arm, this time in the water in the jar. Observe that there is no longer equilibrium. Add water carefully to the beaker on the pan A from a burette until equilibrium is restored. Measure the volume of water added. It is the same as the volume of the cylinder. So the cylinder when immersed in water lost a weight which was equal to the weight of water which occupied the same volume. This experiment has verified a fact which is usually called **The Principle of Archimedes**, viz.: *A solid when immersed in a liquid loses a weight which is equal to the weight of the liquid which would occupy the same volume.*

<sup>1</sup> Griffin's Physical Balance No. 635 is suitable for the purpose, or an ordinary balance with a strip of metal bent into a bridge across the pan. The cylinder can be suspended like a beer barrel by a loop.

<sup>2</sup> A burette may be used in place of the measuring jar.

(iii) **Third Method.**—The principle of Archimedes affords a ready method for determining the relative density of any solid which can be suspended and weighed in water.

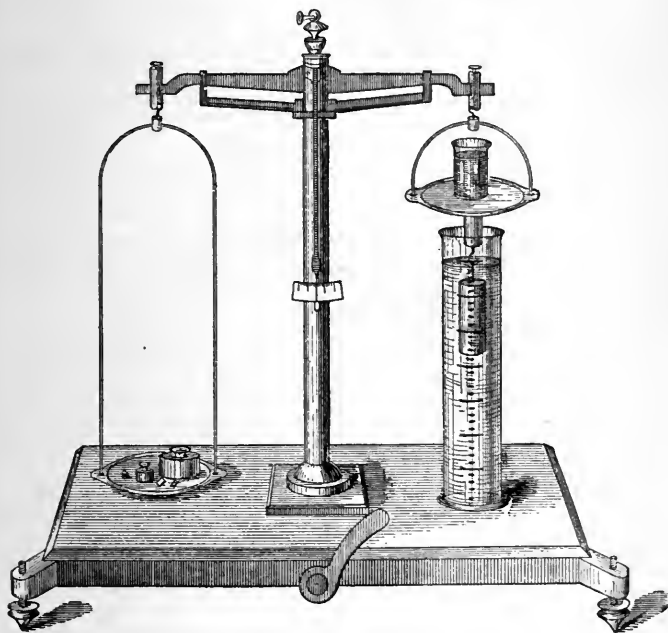


FIG. 27.

**EXPT. 5. To find the Relative Density of a Glass Stopper.**—Suspend a solid glass stopper by a silk thread above the left-hand pan of the balance (Fig. 28). Place a strip of metal (zinc or brass) bent into a bridge across the pan. Place a beaker on the bridge so that the stopper hangs freely within the beaker. Weigh the stopper. Pour water carefully into the beaker until the stopper is completely immersed, and remove any air bubbles which cling to the

stopper with a camel's hair brush. Weigh it again. Thus in an experiment :—

weight of glass stopper in air ... = 15 grams

weight of glass stopper in water = 9 „

Hence,

loss in weight of glass stopper ... = 6 „

Now,

$$\begin{aligned} \text{relative density of glass stopper} &= \frac{\text{weight of glass stopper in air}}{\text{weight of an equal volume of water}} \\ &= \frac{15}{6} \text{ or } 2.5 \end{aligned}$$

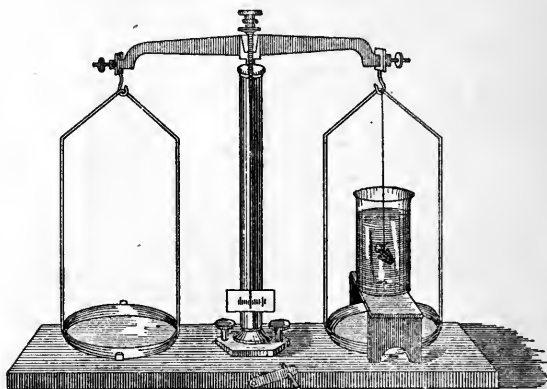


FIG. 28.

### Relative Densities of some Common Substances

Cork ... ..	·25	Tin ... ..	7.3
Wood ... ..	·5 +	Zinc ... ..	7.2
Ice ... ..	·9 +	Iron (cast) ... ..	7.2
Sand ... ..	1.4 +	Copper ... ..	8.9
Marble ... ..	2.7	Silver ... ..	10.5
Aluminium ... ..	2.7	Lead ... ..	11.4
Glass (crown) ... ..	2.5	Gold ... ..	19.3
Sulphur ... ..	2.0	Platinum ... ..	21.5

### EXERCISES

#### Relative Densities of Solids

1. Find by Method (ii) the relative density of—
  - (1) a brass chain.
  - (2) copper nails.
  - (3) glass beads.
  - (4) slate pencils.
  - (5) sand.
  - (6) powdered glass.
  - (7) sulphur.
  - (8) copper wire.
2. Find whether Archimedes' Principle is applicable to—
  - (1) a solid glass stopper.
  - (2) iron nuts or bolts.
  - (3) a leaden cylinder.
  - (4) a silver watch chain.
3. Find by Method (iii) the relative density of—
  - (1) a solid glass stopper.
  - (2) iron nuts, bolts, or screws
  - (3) a silver watch chain.
  - (4) a lump of sulphur.
  - (5) a brass chain.
  - (6) sheet lead.

### EXAMPLES X

#### Relative Densities of Solids

1. Solids, weighing 10, 15, 20 grams respectively, were placed successively in a graduated cylinder partly filled with water, the surface of which read 95 c.c. originally. When the first was added the water rose to 99 c.c., when the second was added it rose to 106 c.c., and when the third was added to 108 c.c. Find the relative densities of the solids.
2. Find the relative densities of the following metals given :—
  - (1) a litre of copper weighs 8950 grams.
  - (2) 75 cc. of zinc weigh 540 grams.
  - (3) 676 cc. of cork weigh 169 grams.
  - (4) 300 cc. of lead weigh 3.42 kilograms.
3. A gold coin weighs 122.5 grains in air, and 115.5 grains in water. Find its relative density.

4. If a silver coin weighs 416 grains, and has the relative density 10·4, what will it weigh in water?

5. Find the loss of weight in water of a 56 lb. cast-iron weight (relative density 7·2).

6. A brass chain weighs 96 grams. When lowered into a measuring jar already containing 100 c.c. of water it raises the water level to 112 c.c. Find its relative density.

7. If 1000 oz. water fill 1 cu. ft., and cast iron is 7·2 times as heavy as water, find the weight of an iron block 4 ins. long, 3 ins. broad, and 1 in. high.

8. A dozen shot made from lead of relative density 11·4 raise the level of water in a measuring jar from 98 c.c. to 122 c.c. Find the weight of each shot.

9. Find the volume of 1 cm. length of platinum wire (relative density 21·5), supposing a piece 1 metre long to weigh 2·15 grams.

10. A 100 c.c. flask filled with water to the mark weighs 125 grams. When 57 grams of lead are added and the water is again adjusted to the mark it weighs 177 grams. Find the relative density of lead.

11. A 100 cc. flask weighs 25 grams when empty. What will it weigh when 30 grams of powdered glass (relative density 3·0) are placed inside, and water is filled in up to the mark?

12. Find the relative density of a mixture of 100 c.c. lead (11·4), and 200 cc. tin (7·3), assuming that they mix without contraction.

13. Find the percentage volume of tin in an alloy of tin (7·3) and copper (8·9), having the relative density 8·7.

14. If a glass stopper weighs 24 grams in air, and 16 grams in spirits of relative density ·8, find the relative density of the glass.

15. A shilling weighs 87 grains, and has the relative density 10·4. What would be the weight of a platinum coin (21·5) of the same size?

16. Find the value of 1 c.c. of silver (relative density 10·5), when silver is worth 2s. 6d. per oz. Reckon 28 grams = 1 oz.

17. A glass stopper weighs 30 grams in air, 17·5 grams in water, and 20 grams in spirits. Find the relative density of the spirits.

## CHAPTER VIII

### THE MEASUREMENT OF THE PRESSURE OF THE ATMOSPHERE

**The Pressure of the Atmosphere.**—We have no difficulty in realising the existence of solids and liquids around us. It is different with the air, and at first it is difficult to believe that there is an invisible substance surrounding us on every hand. Yet the existence of something invisible is at least suggested by the draught from an open window or door, or by the force of a strong wind. If we take a tumbler and force it mouth downwards into water the fact that the water does not completely fill it shows at least that the tumbler contains something which opposes the entrance of the water. Such facts help us to realise that we live in an ocean of air, which extends in all probability to a distance of at least 200 miles from the earth.

We are aware of the pressure of the air in a strong wind, but in a room in which the air is at rest we are not conscious of any pressure.

**EXPT. I. To show that Air is exerting a Pressure even when we cannot feel it.**—Bend a glass tube as in Fig. 29. Pour in some coloured water. The water stands at exactly the same height in both arms. Why? Because any pressure of the air is exerted equally on both surfaces. If now we attach a piece of rubber-tubing to the arm A B, and suck out some of the air from that arm, what may occur? Try. The water rises in the arm A B, and falls in the other. as repre-

sented in Fig. 30. Secondly, blow in a little air. The water falls in the one arm, and rises in the other.

The rise and fall of the water in the arm A B shows that the air was originally exerting a pressure equally upon the surface of the liquid in both arms; and that the pressure of the air in A B was diminished when air was sucked out, and increased when air was blown into it.

EXPT. 2. Place one end of a glass tube below the surface of some water. Apply the mouth at the other end and suck gently. As the air is gradually inhaled, the water rises in the tube.



FIG. 29.

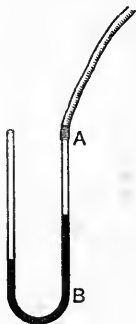


FIG. 30.

This suggests that the air is exerting a pressure on the surface of the water outside the tube, so that the water is forced up the tube as soon as the air is partially removed from the inside, and when the pressure there is consequently diminished.

EXPT. 3. Fit a thistle-funnel in one end of a piece of thick rubber pump-tubing. Fix the other end to an air-pump. Moisten the palm of the hand and press the mouth of the

funnel against it. Then proceed to exhaust the air from the tubing. Observe that the funnel adheres firmly to the hand.

It is clear, therefore, that the air even when quite still is exerting pressure, and the fact that no force may be felt upon the extended hand arises in the first place from the pressure being exerted above and below it equally. Then why is the hand not squeezed flat? Because there is an internal pressure of the blood in the veins and arteries and other fluids within the body, greater than that of the atmosphere.

Torricelli, an Italian physicist, first showed in 1643 in the following way how to measure the pressure which the atmosphere is always exerting.

EXPT. 4. **To make a Barometer.**—Take a thick-walled glass tube, about 4 mm. internal diameter, and 90 cm in length.



Seal one end in the blow-pipe flame. Holding the open end upwards, fill the tube completely by pouring clean dry mercury into it through a small funnel. Gently tap the tube so as to dislodge air bubbles. Pass a bubble the length of the tube so as to sweep out smaller bubbles. Fill the tube completely with mercury. Close the open end firmly with the thumb, and invert the tube in a mortar containing mercury (Fig. 31). The mercury sinks in the tube leaving a clear space A B at the top of the tube. Tilt the tube: if the mercury does not rise completely to the top, some air bubbles have not been removed; these must then be displaced by inverting the tube again, and gently tapping. Measure the length of the mercurial column above the level of the mercury in the mortar. It is about 76 cm., or 30 inches. From the way in which the tube was set up it seems impossible that the space A B can contain any air.

**EXPT. 5. To show that A B is a Vacuum.**—(1) Tilt the tube gradually, keeping the unsealed end still below the mercury in the mortar. The mercury rises in the tube, and finally completely fills it.

(2) Pour some water above the mercury in the mortar. Raise the tube carefully so that the open end C lies in the water. The mercury at once runs down, and water not only takes its place, but completely fills the tube.

It is evident, therefore, that the space A B is free from air. After the author of the experiment this space is usually called the Torricellian vacuum.

From this and the former experiments it appears that the column of mercury B C must be supported by the pressure of the air on the mercury in the mortar.

Robert Boyle, who invented the air-pump, saw that a crucial

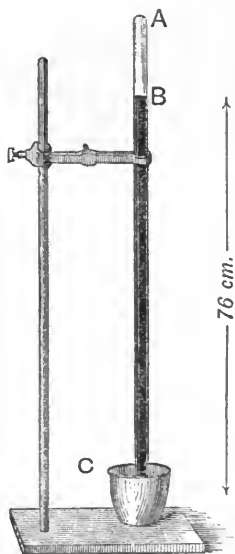


FIG. 31.

experiment<sup>1</sup> to test the truth of this view could be made by placing the whole apparatus used in Expt. 4 within the receiver of the air-pump. He found to his delight that the mercury fell further and further in the tube as the air was gradually exhausted, until it was only an inch or two above the level of the mercury outside. On re-admitting air into the receiver he found that the mercury rose to its original level in the tube.

It is evident, therefore, that the *height of the mercury in the tube above the level outside is a measure of the pressure of the atmosphere*. The instrument was called by Robert Boyle a *barometer*, and ever since his time it has been in use for this purpose.

**Variations in the Atmospheric Pressure.**—If we imagine the earth to be enclosed in a vast ocean of air, we should expect the pressure of the air to become less and less as we ascend above sea-level, and this should be made evident by our barometer. It was shown to be so by Périer in 1648. He ascended the Puy-de-Dôme, in the Auvergne Mountains of France, and found that the mercury fell about 3 inches. The barometer even shows that in the course of any railway journey there are variations in the atmospheric pressure according to the height above sea-level.

Further, at any one place, for example at Manchester, the barometer shows that variations occur from time to time in the atmospheric pressure. These variations arise from the movements which are continually taking place in our atmosphere. Sometimes in a storm the height of the barometer will change by as much as 1 inch within 24 hours.

Apart from such irregular changes it is remarkable that daily variations can also be detected. The pressure tends to be greatest at about 9 o'clock in the morning and evening, and least at about 3 o'clock in the morning and afternoon.

**The Standard Atmospheric Pressure.**—The *average* pressure, however, at any place on the sea-level is represented by about 30 inches or 760 mm. of mercury. It is customary, moreover, to regard the pressure which is measured by 760 mm. of mercury as the standard atmospheric pressure. In other

<sup>1</sup> Boyle's experiment should be shown as a Lecture Experiment. A strong filter-flask, fitted with a barometer tube, and an exhausting air-syringe attached to the side tube of the flask may be employed.

words, the pressure exerted by a column of mercury 760 mm. in length is called a pressure of *one atmosphere*.

It is useful to remember that this pressure is about 14·7 lbs. weight upon every square inch, or  $760 \times 13\cdot6$  or 1033·6 grams weight upon every square centimetre.

**The Fortin Barometer.** — The variations in atmospheric pressure which occur at any one place are often so slight that they could not be detected with the simple form of barometer described in Expt. 4. It is necessary to employ devices which enable us to read the length of the mercurial column with great nicety. Perhaps the most perfect form of barometer is that known as the Fortin instrument.

The peculiar feature of this barometer is the *cistern*, which is shown in section in Fig. 32. The air in the cistern is in communication with the outside air through a piece of porous chamois leather at the top of the cistern. The upper portion of the wall of the cistern is of glass, through which the mercury level can be observed. The lower portion of the inner wall is of boxwood, and to this is fastened a bag of leather, which supports the mercury. Fig. 32 shows that the bottom of the bag can be raised or lowered by a screw. Hence, by turning this screw it is always possible to bring the surface of the mercury in the cistern just into contact with the tip of the ivory pointer which is fixed vertically above the mercury. Whatever the variation in the pressure of the air, a constant level can, therefore, be ensured in the cistern, from which to measure the height of the mercurial column. The level of the top alone of the mercurial column will vary, and it will only be necessary to read its

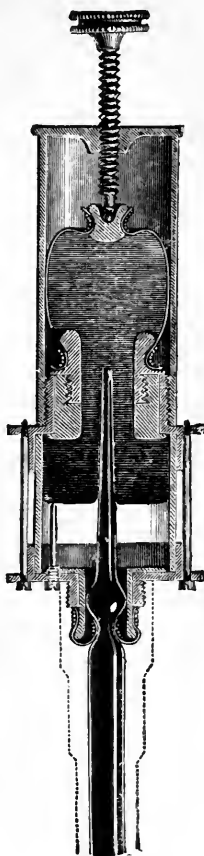


FIG. 32.

position. This is read accurately by means of a scale and vernier (see p. 20).

**The Method of Reading the Barometer.**—The mode of taking an observation is this:—Adjust the mercury in the cistern by turning the screw at the bottom of the cistern, so that the ivory point is *just* brought into contact with the surface of the mercury, but does not depress it; the ivory point and its reflected image in the mercury should appear to just touch each other and form a double cone. Then gently tap the tube with the finger, to prevent the mercury adhering to the glass. Next adjust the vernier so that its two lower edges shall form a tangent to the *convex* surface of the mercury. In fact the front and back edges of the vernier, the *top* of the mercury and the eye of the observer must be in the same straight line. Read the vernier. After the reading has been taken, the mercury in the cistern should be lowered by turning the screw at the bottom until the surface is well below the ivory point, otherwise dirt will collect and the mercury will oxidise immediately underneath the ivory point, and its reflected image will become indistinct.

#### EXERCISES

1. Attach a U-tube containing water by means of a rubber-tube to the coal-gas supply. What do you observe? Can the pressure of the gas be measured?
2. Repeat Ex. 1, using another tube of either wider or narrower bore. What effect do you conclude that the bore of the tube has upon the result?
3. Repeat Ex. 1 and 2, but using methylated spirits in place of water.

## CHAPTER IX

### FITTING UP APPARATUS

NOTHING is more frequently needed in any laboratory<sup>1</sup> than some source of heat. If a chemist had no fire or flame at his service he would be able to do very little. The alchemists generally used some form of furnace. In quite modern times<sup>2</sup> coal gas has come into common use, and is supplied in every laboratory. Is an ordinary bright coal gas flame suitable for laboratory purposes?

Hold a piece of white cardboard for a few moments across a bright gas flame. It becomes very sooty. If, then, a glass vessel were heated in such a flame it would quickly become coated with soot, and it would be difficult to watch the progress of what might be going on inside it. But it is frequently most important to be able to see the contents of a vessel. Consequently, a bright coal gas flame is not a satisfactory source of heat, and it would be a very great gain to have a flame which would not deposit a black curtain of soot upon any vessel heated by it. A burner which would give such a flame was devised by Robert Wilhelm Bunsen, Professor of Chemistry in the University of Heidelberg. A great teacher and a famous investigator, he has now withdrawn from active work, and is living in retirement in his old University city on the banks of the Neckar.

**The Bunsen Burner.**—If air be allowed to mix with coal gas before it is burnt, a blue flame hotter and smaller than the

<sup>1</sup> See Appendix on Laboratory Fittings, and Lists of Apparatus and Chemicals.

<sup>2</sup> The streets of London were first lighted with coal gas in 1812.

ordinary luminous gas flame results. Examine a Bunsen Burner. At the bottom of the tube A (Fig. 33) is a collar B. Oval windows or holes are pierced in the tube and in the collar. Either the tube or the collar can rotate so that these windows can be set opposite to each other, admitting air or alternately excluding air.

(1). Close the holes. Light the gas. The flame is yellow.

Bring a piece of white cardboard into the flame; it becomes sooty.

(2). Open the holes.

The flame becomes blue. Hold a piece of cardboard in it. It becomes brown and singed, but not sooty. Hold a match across the flame about one inch above the tube. The wood is burnt in two places, showing that there is an inner cooler region. Try whether you can introduce a match head into the middle of the flame without lighting it.

The inner tube C of the burner will occasionally get blocked by molten substances falling on it.

It may be cleaned by washing with water or by picking it out with a pin.

(3). Light the burner with a blue flame. The tube contains a mixture of gas and air. Press gently upon the rubber gas supply tube, making the flame smaller and smaller. Suddenly the flame will shoot down the tube and it will be seen *burning below*. Notice the curious smell. Light the burner at the top. The flame is yellow although the holes are open. Any attempt to turn the tube will now lead to burnt fingers, since the inner flame makes the tube very hot. It is better to turn the gas out, and then light it again.

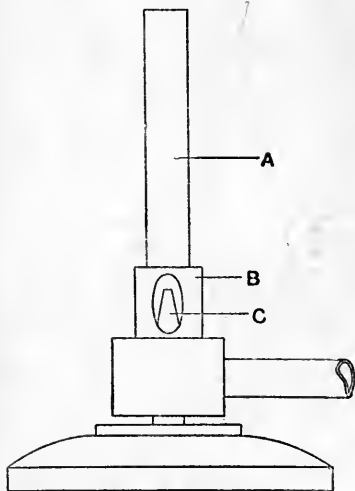


FIG. 33.

When a very small flame is required the supply of air should be reduced as well as the coal gas.

**Glass** is useful to the chemist on account of (1) its transparency, (2) the ease with which it can be worked into any shape when hot and plastic, (3) its not being attacked by most chemicals. Its disadvantage is its brittleness. It is particularly liable to break on heating, because it is a bad conductor of heat, and therefore one part of the glass expands on being heated before neighbouring parts become warm.

**To Illustrate the Liability of Glass to Crack when Heated.**—(1) Hold a piece of glass rod an inch long in the flame of the foot-blowpipe. The glass cracks before the fingers feel the heat.

(2) Warm another piece of glass rod very gradually. Then pour cold water on it. It cracks.

(3) Fill a test-tube one-third full of water. Hold it in the flame so that the flame plays on the glass near the water line. The tube cracks.

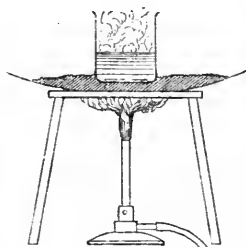


FIG. 34.

**To Boil Water in a Glass Vessel.**—(1) Try again to warm water in a test-tube, as follows. Let

the flame play on the bottom of the test-tube. Turn the tube round between finger and thumb. Shake the tube gently so as to keep the water dancing up and down the sides of the tube, and so warm them gradually. The water boils without accident.

(2) Place a piece of iron wire gauze on a tripod-stand, or on a ring of a retort-stand. Place a glass beaker or flask containing some water on the gauze. Put a Bunsen flame under the gauze. The gauze distributes the heat over the bottom of the vessel, and the water can be heated until it boils without accident.

(3) Place a little sand in an iron tray, and put the tray upon a tripod-stand (Fig. 34.) Place a beaker containing some water upon the sand, and heat the tray with a Bunsen burner. The heat is distributed by the sand, so that the beaker can be heated safely until the water boils.

The tray containing the sand is called a *sand-bath*.

### To Make a Wash-Bottle

It will generally be possible to fit up for ourselves the apparatus required for our experiments. We shall require glass flasks, corks, glass tubing, and india-rubber tubing, and some knowledge of how to use them.

Suppose we learn in the first place how to make the apparatus shown in Figure 35. This is called a *wash-bottle*. It is useful for sending a stream of water in any desired direction.

By blowing at *d* a jet of water spurts from *c*, and the india-rubber connection *e* allows of the water being sent in any direction. We shall require a flask, a cork, about half a yard of glass tubing, and a small piece of rubber tubing.

The cork (wooden) selected must be just too large to fit into the mouth of the flask. It can be made to fit the neck by rolling and squeezing it under the foot. Begin gently, turning it round and round, and press harder and harder. The cork now fits well into the neck of the flask, and being somewhat elastic presses tightly against the glass all round.

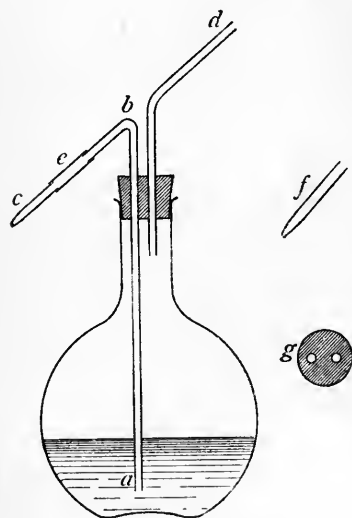


FIG. 35.

An india-rubber stopper may be used instead. It should fit the mouth easily, since it will be enlarged when the glass tubes are passed through it.

**To bore a Cork.**—A hollow brass tube sharpened at one end is used, called a cork-borer. Choose one a little narrower than the glass tubes which are to go through the cork. At the narrow end of the cork put marks where the holes are to be made. Stand the cork on the table, narrow end up. Keep the borer care-



fully upright. Drive it gently into the cork with a screw motion, half a turn at a time, pausing to see that it keeps upright. Along with a set of cork-borers is supplied a little solid brass rod, which can be used both to press the core of cork out of the borer and as a handle to turn the borer round. (The borers will require sharpening periodically.)

The holes so made may be found a little too small. If so, enlarge them carefully with a rat-tail file. Roll the cork whilst filing so that the hole may be enlarged evenly and remain round. These files can be used for boring as well as for enlarging holes, but care must be taken to keep the holes quite round.

**To cut Glass Tubes.**—Take a sharp triangular file. Make a nick in the glass tube about 15 cm. from one end. Hold the tube in both hands. The two forefingers should be together immediately underneath the nick. The two thumbs are above. Turn the wrists outwards and downwards, round the forefingers as axes, at the same time pulling the hands apart (Fig. 36). The glass cracks clean across.

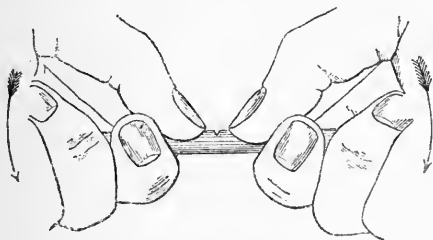


FIG. 36.

**To bend Glass Tubing.**—Hold a piece of glass tubing lengthwise along a *flat* gas flame, as shown in Fig. 37 (fishtail or batwing, not the Bunsen flame). Keep it moving slowly round and round and also to and fro. As soon as it feels slightly soft, hold it by one hand alone, and cease rotating it. The free end will gradually fall (Fig. 38). As soon as it is bent to the required angle remove it from the flame. A considerable length has been evenly heated, and a smooth bend has resulted. Now

cut two lengths of 15 cm. and 25 cm., and bend them at angles of  $135^\circ$  and  $45^\circ$  respectively as shown in Fig. 35.

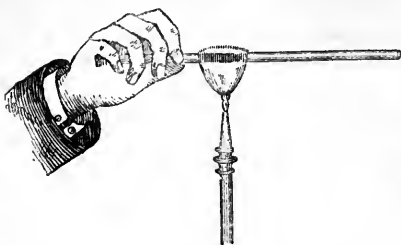


FIG. 37.



FIG. 38.

**Mistakes to avoid in bending Glass.** (1) Hold a piece of glass tubing across the blue Bunsen flame. It becomes hot in two places. When soft, bend it. The glass is found creased at the bend. (2) Hold a piece of glass tubing across an ordinary flat gas flame. The tube is heated this time in one place. Bend it. The tube is flattened on the outside of the bend and thickened on the inside. When cold, try to bend the tube. It snaps easily at the bend.

**To round the Ends of Glass Tubing.**—The sharp edges left at the ends of the glass tubing must be rounded, otherwise the tubes will not slide easily through corks, and are apt to cut any rubber tubes that are slipped over them. Round the ends of the tubes just made (*a*) by grinding the sharp edge on a file or on a stone floor, (*b*) by holding the end in a flame until the glass just melts at its edges.

**To round the Ends of Glass Tubing with the Mouth Blowpipe.** A *Blowpipe* (Fig. 39) is used to turn a sharply

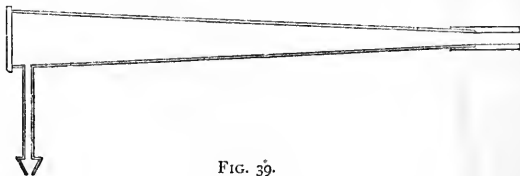


FIG. 39.

pointed flame in any required direction. Use the yellow flame of the Bunsen burner. Turn it down until it is two or three inches

high, and just does not flicker. Point the blowpipe jet a little downwards and hold it just above and to the side of the burner as shown in Fig. 1. Blow gently so that the flame just fails to hiss. Keep up a steady current of air for as long as you can. The cheeks soon empty, but may be refilled and kept distended by moving the tongue in the roof of the mouth, and inhaling through the nose. Blow from the cheeks, not from the lungs. A glass mouthpiece will make the blowpipe easier to the lips. The student will soon learn to maintain a steady stream of air for five minutes without stopping.

Cut a piece of glass tubing 5 cm. long. Round both ends. Do the same with a piece of glass rod 4 cm. long, but warm the rod gently before heating it with the blowpipe.

**To draw out Glass Tubing.—**

(1). Holding a piece of glass tube in both hands, heat it at one place in the blue Bunsen flame. Remove it when soft from the flame and pull.



FIG. 40.

The tube is drawn a little at A (Fig. 40). At B it is pulled out much finer, because it happens to be hotter, and it has broken there because the pull was too hard. The lump C was in the cooler central portion of the flame, and never became properly heated.

(2). Heat another tube in a flat flame, holding it lengthwise and turning it slowly round and round. Pull the two ends



FIG. 41.

apart. There is no lump, because the tube was heated evenly, but the finely drawn out tube has softened at once, and is bent down (Fig. 41).

(3). Heat another tube as in Expt. 2 until the tube feels

limp in the fingers. *Remove it from the flame.* Draw slowly at first, then more quickly, but gently rotating each end until



FIG. 42.

the hands are wide apart. Hold it rigidly for a few seconds, until the glass stiffens (Fig. 42). The glass is now drawn out perhaps a yard long or more. Cut the tube with a very sharp file gently at A and B. Notice how elastic this section A B is. It can be bent nearly to a circle. If bent too far it snaps, and the pieces spring back straight again. Dip the end B into water and blow at A. Bubbles arise. Dip another piece of this fine tube into an inkpot and note the result. Then, A B is really hollow all through. The ends A C can be used for the wash-bottle jet. If B D is rather longer it will do for mounting platinum wire.

The foot blowpipe flame is better than a flat flame for drawing out glass tubing.

**To narrow or close an End of a Glass Tube.**—Incline

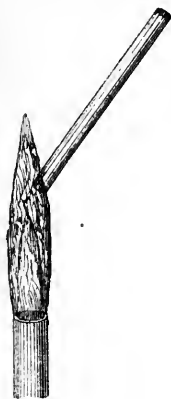


FIG. 43.

a glass tube 20 cm. long downwards at an angle of  $60^\circ$ , and hold the lower end in the edge of a Bunsen flame at a point about two-thirds of the height of the flame as shown in Fig. 43. The edge becomes red hot, fuses, and gradually contracts. Remove the tube from the flame before it is quite sealed up. Cut the glass about 4 cm. from the end. This portion can be used for the wash-bottle jet.

Hold one end of the remainder of the tube for a longer time in the flame and close the end entirely.

**Rubber tubing** sometimes becomes stiff when not in use. Warm gently and pull it about; it soon softens. Cut off a piece an inch and a half long for future use.

If a rubber tube is too wide for a glass tube it can be made to fit by doubling it back on itself.

**To fit together the Wash-Bottle.**—Moisten the end a

of the long tube  $a\ b$  (Fig. 35). Push the long tube through the cork with a zig-zag screw motion. Hold the tube near the cork and avoid holding it at the bend  $b$ , where it might break and cause a nasty cut in the palm of the hand. Fit  $d$  in the cork also. Now fit the cork into the bottle. The end  $a$  should come nearly to the bottom of the flask, but not so near as to risk touching it.

**To test whether the Flask is Air-tight.**—Close the glass tube at  $e$  and blow at  $d$ . Does the air hiss out through the cork? Again closing  $e$  suck at  $d$ . The tongue is sucked back and held at the end of the tube. Attach the jet  $c$  by means of the rubber tubing. If the flask is now filled with water it is ready for use.

#### **To use the Wash-Bottle.-**

- (1) Blow at  $d$ . A fine jet of water issues from  $c$ .
- (2) Invert the bottle. A large stream of water issues from  $d$ . It is often well to keep the bottle filled with distilled water. If a supply of hot water is required, the bottle can be heated over a flame.

LABORATORY RULES, if not provided by a teacher, should be made by the student for himself:—

1. Clean apparatus, and put it away in its proper place as soon as done with.
2. Matches, filter-papers and rubbish must not be thrown into the sink, but put into the refuse box.
3. See that gas and water are turned off before leaving the room.

### EXERCISES

#### **Working of Glass. Fitting up Apparatus**

1. Cut 5 tubes 5 cm. long, 5 tubes 15 cm. long, 5 tubes 30 cm. long.
2. Prepare tubes bent as follows:—
  - 3 tubes, 15 cm. long bent in the middle at right angles.
  - 3 tubes, 30 cm. long bent at right angles near one end.
  - 2 tubes, 15 cm. long bent at  $120^\circ$  in the middle.
  - 2 tubes, 30 cm. long bent at  $60^\circ$  near one end.
  - 2 tubes, 40 cm. long to be bent at right angles near one end. Cut off the shorter limb  $\frac{1}{2}$  inch from the bend.

2 tubes, 25 cm. long to be drawn out near one end and cut in two. The shorter pieces can serve as wash-bottle jets, the longer ones can be used as pipettes.

1 tube, 40 cm. long bent into an equilateral triangle.



FIG. 44.

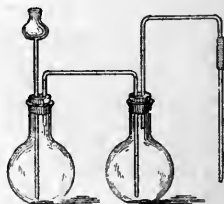


FIG. 45.

3. Practice pulling out glass tubes to full arms' length. Break up the capillary tube so formed into 6 inch lengths, and save them.

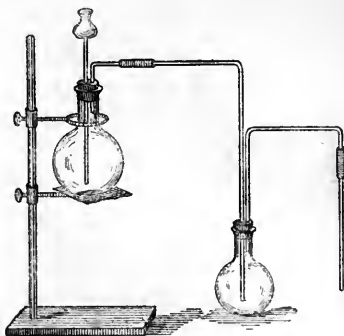


FIG. 46.

4. Cut 4 pieces of glass tubing 5 cm. long. Round both ends of each. Do the same with 4 pieces of glass rod, each 4 cm. long.

5. Fit up the pieces of apparatus shown in Figs. 44, 45, and 46.

## CHAPTER X

### CHANGE OF STATE—LIQUEFACTION AND SOLIDIFICATION

WE are acquainted with many different solids and liquids.

What is the effect of heat upon them?

**Many Solids are changed into Liquids by Heat, which solidify again on cooling.**

EXPT. 1. Place some powdered *sulphur* in a test-tube. Heat very gently in a Bunsen flame. The sulphur melts, and forms a yellow liquid. Place the test-tube in some cold water. The liquid quickly solidifies again, forming a solid cake of sulphur.

EXPT. 2. Place a lump of *paraffin wax* in a test-tube, and heat it gently in a flame. The wax melts, and forms a clear liquid. Cool the test-tube. The liquid is quickly changed into an opaque white solid mass.

EXPT. 3. Place a few pieces of *tin* in an iron spoon. Heat in a good Bunsen flame. The tin melts completely. Skim off the earth-like dross on the surface of the molten metal with an iron wire. Pour the liquid in a thin stream into a bucket of water. It is converted into a bright crumpled mass of the solid metal again

### EXERCISES

Find the effect of heat and subsequent cooling upon :—

- |                                   |                            |
|-----------------------------------|----------------------------|
| (1) sealing wax.                  | (6) lead.                  |
| (2) shellac.                      | (7) zinc.                  |
| (3) vaseline.                     | (8) soda.                  |
| (4) bees' wax.                    | (9) washing-soda crystals. |
| (5) material of a stearin candle. |                            |

Let us study one particular case more fully, so as to obtain further information as to what occurs when a solid is converted by heat into a liquid.

**EXPT. 4. The Effect of Heat upon Ice.**—Break some ice into small fragments, the smaller the better. Pack them into the lower half of a beaker. Plunge a thermometer into the mass. The temperature is  $0^{\circ}\text{C}$ . Warm the beaker on a sand-

bath over a very small flame, and stir the mass continually with the thermometer, occasionally removing the beaker for a few seconds. Read the thermometer frequently. The ice rapidly melts. Although so much heat is added, nevertheless the temperature remains fixed at  $0^{\circ}$  to  $1^{\circ}\text{C}$ ., until the whole of the ice is melted. Only then does the temperature begin to rise and continue to do so.

**The Melting Points of Solids.**—It is evident that while ice is being liquefied no rise of temperature occurs. In other words, the change of state takes place at a definite temperature. This temperature is called the *melting point* of the ice.

Other substances behave similarly to ice, and we may now proceed to determine in a simple way the melting points of some of them.

**EXPT. 5. To find the Melting Point of Paraffin Wax.**—Take a capillary glass tube. Close one end by holding it in a flame. Introduce some small particles of paraffin wax to a depth of about 0.5 cm. Attach it to a thermometer by means of a thin slice of rubber tubing so that the wax is opposite the middle of the bulb. Support a beaker, of about 100 c.c. capacity, half-filled with water, over wire gauze on a ring of a retort-stand. Place a glass stirrer *b*

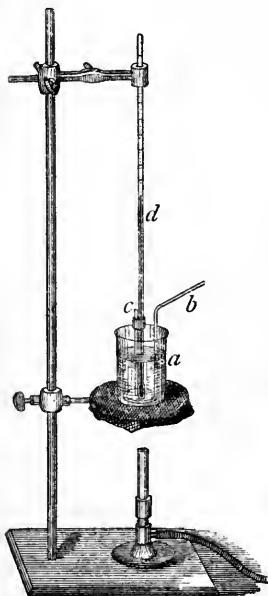


FIG. 47.



inside, the end of which has been bent into a circle at right-angles to the stem. Support the thermometer so that its bulb is in the centre of the water (Fig. 47). Heat gently, stirring constantly. As soon as the paraffin wax melts read the temperature. Probably you have been too late, and the reading is too high.

Allow the water to cool slightly. Take a fresh capillary tube containing paraffin wax and ascertain the *exact* melting point, removing the flame from time to time and stirring thoroughly, so that the temperature rises very slowly. Read to the tenth of a degree.

### EXERCISES<sup>1</sup>

Find the melting points of (1) iodine, (2) naphthalene, (3) filings of Wood's metal, (4) stearin candle material, (5) sulphur.

We have found that different solids melt at different temperatures. Usually, moreover, the melting point is in each case very definite. Pure sulphur, for example, has a well-defined melting point. Impure substances, on the other hand, do not melt entirely at one temperature but gradually through a range of temperature. Hence use may be made of melting points in order to distinguish and identify substances, and to detect the admixture of foreign ingredients. The melting point is indeed one of the best defined of physical "constants." Here are a few melting points, many of which, however, have been determined by other and special methods.

TABLE OF MELTING POINTS.

Gold	... ..	1,062° C.	Tin	... ..	233° C.
Silver	.. ...	961°	Sulphur	.. ...	115°
Borax	... ..	878°	Paraffin, about	...	54°
Common salt...		851°	Butter, about	...	33°
Zinc	... ..	419°	Ice	... ..	0°
Lead	... ..	326°	Mercury	... ..	-39°

**Latent Heat of Liquefaction.**—We found that while ice was being melted no rise of temperature occurred, in spite of

<sup>1</sup> It will be found necessary to use strong sulphuric acid in place of water in the case of iodine and sulphur.

the addition of so much heat. In a similar way heat apparently disappears while any other solid is being changed into a liquid. This heat which alters the state of a solid, but does not raise its temperature, and which therefore apparently disappears, was first investigated by Dr. Black of Edinburgh about the year 1760. The question may be asked, whether when such a liquid is reconverted into the solid state the heat will become sensible again.

**EXPT. 6. To show that Heat is given out when a Liquid solidifies.**—Take some crystals of a substance known to the photographer as “hypo,” and to the chemist as sodium thiosulphate. Put them in a perfectly clean flask. Heat gently, until the crystals are completely melted. Introduce a thermometer. Place a plug of cotton-wool in the neck of the flask. Allow to cool in a quiet place. Then remove the plug, and introduce a small crystal of sodium thiosulphate. The liquid quickly solidifies. At the same time the thermometer rises many degrees, and the flask becomes quite hot.

It is evident then that the heat, which apparently disappears when a solid is liquefied, can be recovered if the liquid be again caused to solidify. Hence the heat is said to be hidden or latent, and it is usually spoken of as *latent heat of liquefaction*.

**EXPT. 7. To show that Heat is taken up when Sal-ammoniac is dissolved in Water.**—Powder 50 grams of sal-ammoniac. Place 100 cc. of cold water in a beaker. Read its temperature. Add the sal-ammoniac. The beaker feels chilled, and dew collects on its outer surface. Find the fall in temperature. There may then be an absorption of heat when a substance is *dissolved* in water.

### EXERCISES

(1) Repeat Experiment 6 with crystals of sodium acetate, ammonium nitrate, or sodium sulphate, dissolved in one quarter their weight of water.

(2) Find whether there is any change of temperature when sal-ammoniac, salt, washing-soda, and nitre are respectively dissolved in cold water.

**Freezing Mixtures.**—Advantage is taken of the fact that heat becomes latent in the liquefaction of any substance to produce “freezing mixtures.” For, supposing the liquefaction of a substance is induced to take place, and yet no heat is supplied from any external source, the heat which must be absorbed in the process of liquefaction will be taken out of the substance itself or out of anything in contact with it, and a considerable fall of temperature may ensue. For instance, if 1 part of common salt is mixed intimately with 4 parts of pounded ice or snow the mass liquefies because the freezing-point of *brine* is far below that of water. Hence, at the temperature of the experiment, the mixed materials ought to be liquid. But in order to liquefy they must absorb heat. They, therefore, sacrifice their own temperature in order to get this heat, and the temperature sinks to about  $-22^{\circ}\text{C}$ .

#### PROBLEMS

Find whether freezing mixtures can be made from—

- (1) Glauber's salt and strong hydrochloric acid.
- (2) Ice and crystallised chloride of calcium.

## CHAPTER XI

### CHANGE OF STATE—VAPORISATION AND CONDENSATION

WE have found that many solids are changed into liquids by heat. What is the effect of heat upon liquids?

**Many Liquids are changed into Vapours by Heat, which liquefy again or solidify on cooling.**

EXPT. 1. Place a few crystals of *iodine* in a test-tube. Heat very cautiously. The iodine first melts, and then violet vapours are given off. These are condensed again on the upper and cooler portion of the tube. Examine it carefully. Crystals of iodine have been formed.

The violet vapour may also be condensed in the cooler air and form a mist consisting of small liquid drops, which then solidify.

EXPT. 2. Place a drop of *mercury* in a test-tube. Heat carefully in a flame. A mirror of minute globules of mercury is formed upon the upper and cooler portion of the tube, although there is no visible vapour.

We know also that water when heated boils at  $100^{\circ}$  and is converted into steam, which is readily condensed as a liquid again.

EXPT. 3. **To show that Steam is an Invisible Vapour.**—Boil some water in a flask. A white cloud appears at the mouth of the flask, but nothing is seen in the flask above the water. The white cloud consists of tiny drops of water—"water-dust," it has been called.

In all these experiments liquids are found to change into

vapours and then on cooling into the same liquids again. Such changes, and those in the case of the melting of solids, are instances of what are called *physical changes*.

Let us repeat the vaporisation of water in another way.

**EXPT. 4. The Distillation of Water.**—(1) Take an 8 oz. flask *a*, and half fill it with water. Introduce a few pieces of broken glass or of clay pipe. Bend one end of a long piece of glass tubing as shown. Fit the tube by means of a cork, through which a thermometer may also be passed, to the flask.

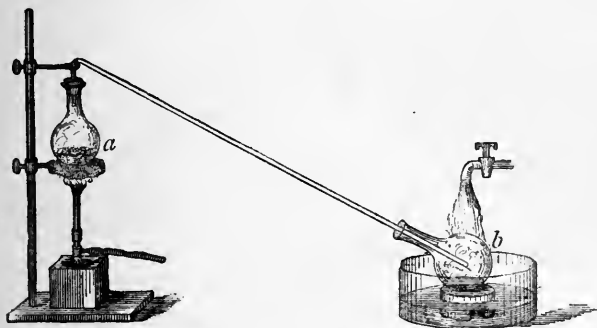


FIG. 48.

The bulb of the thermometer should be above the surface of the water. Support the flask on wire gauze upon the ring of a retort-stand. Place the other end of the glass tube within a flask, *b*, underneath a tap of cold water (Fig. 48). Heat the flask. As soon as the water has begun to boil, the steam passing down the tube is condensed on its cold surface, or, if the glass tube becomes very hot, in the other flask. The condensation of the steam, and the collection of the water produced is spoken of as *distillation*.

(2) The following is a better apparatus for the distillation of water. Take a flask and fix in its neck by means of a cork a thermometer and a bent glass tube as shown in Fig. 49. Take also a "Liebig's condenser," which consists of a long narrow tube, the central portion of which is surrounded by a wider tube, so that cold water can be circulated between them. Select a cork which fits the tube of the condenser. Fix the cork on the side tube of the flask. Place some water and a piece of clay

pipe in the flask. Fix a thermometer in the neck of the flask by means of another cork, so that the bulb is just below the point of junction of the side tube.

Then, supporting the flask upon wire gauze, connect it with the condenser. Circulate cold water through the condenser. Heat the flask. When the water has begun to boil, the vapour, passing the bulb of the thermometer and down the tube, is

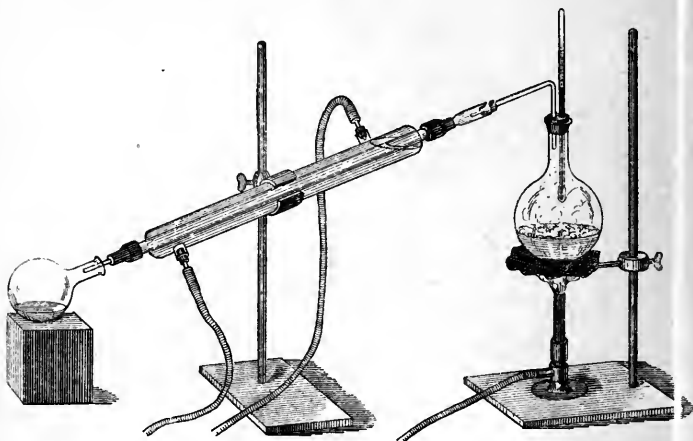


FIG. 49.

completely condensed on the cool surface of the tube. Collect in a beaker the water thus condensed. Meanwhile read the thermometer. It stands constantly at  $100^{\circ}$ .

**The Boiling Points of Liquids.**—We have just found that, while water is being boiled, no rise of temperature takes place, the change of state from water to steam occurring at a definite temperature, called the *boiling point* of the water.

Is the same true of other liquids? Do they all possess the same boiling point?

#### EXERCISES

Find the boiling points of (1) turpentine ; (2) benzene ; (3) alcohol ; (4) toluene, using the apparatus shown in Fig. 49.

Different liquids are thus found to boil at different temperatures. Moreover, when a liquid is pure the boiling point is usually very definite and readily determined. Hence, it is a most useful physical constant, and should be noted if possible in the case of every pure liquid.

TABLE OF BOILING POINTS.

Sulphur	...	...	440° C.	Benzene	...	...	80°·4 C.
Mercury	...	...	357°	Alcohol	...	...	78°·3
Toluene	...	...	110°·3	Chloroform	...	...	60°·2
Water	...	...	100°	Ether	...	...	35°

**Non-Volatile Liquids.**—Some liquids cannot be distilled or converted into vapours which will yield the same liquids again when cooled.

**EXPT. 5. Olive Oil cannot be distilled.**—Half fill a small retort with olive oil. Fit a thermometer in the tubulure, so that the bulb is in the oil. Introduce the neck of the retort into a small flask. Heat the retort over wire gauze. Observe that the temperature rises continually higher and higher. As soon as 300° C. is reached, remove the thermometer, and replace the stopper of the retort. Heat further. After a time vapour is at last evolved, and being condensed in the neck, a few drops of liquid are collected in the flask. Then remove the flame. Notice the odour of the liquid in the receiver. It is very acrid, and clearly not that of unchanged olive oil. Indeed, this liquid is not olive oil at all, but another substance produced by the heat breaking up the oil.

Liquids such as olive oil, which cannot be distilled, are said to be *non-volatile*.

## CHAPTER XII

### CHANGE OF STATE. THE EVAPORATION OF WATER AND OF OTHER LIQUIDS

**The Evaporation of Water.**—We are often, perhaps, inclined to suppose that it is only when water is boiled that it is converted into vapour. Yet this is far from being true. Wet clothes become dry when hung in the open air, particularly on a windy day, and the puddles in the road after rain are quickly dried up in the sunshine. The quiet and invisible passage of a substance from the liquid to the gaseous state at all temperatures is called *evaporation*.

So large a part of the surface of the globe is covered with water in sea, river, or lake, that an enormous amount of evaporation is always going on. It is evident, therefore, that the air must always contain a very large amount of aqueous vapour.

But is there any limit to the amount of aqueous vapour which the air may contain, and what conditions affect the rate at which evaporation proceeds? We must now search for answers to these questions.

Valuable information about the evaporation of liquids can be obtained by introducing them into the Torricellian vacuum of a barometer tube.

**EXPT. I. The Evaporation of Water in a Vacuum.**—Fill two dry glass tubes A and B, each sealed at one end, with warm dry mercury, and invert them in a trough of



mercury (Fig. 50). Introduce a drop or two of water into B under the mercury by means of a pipette C with a curved end. The water being lighter than the mercury rises to the surface in the tube. The mercury column at once falls a considerable distance. This cannot be due to the weight of the drop of water. Observe closely the top of the mercury column. The drop of water has nearly or entirely disappeared. The water has then evaporated very rapidly in the vacuum. The depression of the mercury must be due to a *pressure* which the invisible aqueous vapour is exerting, in the same way that air exerts a pressure.

Supposing we introduce some more water, will the mercury sink further? We must test this by experiment. Add a few more drops of water so that a little water rests upon the top of the mercury column. The mercury does not sink any further after this, even though two or three drops of water are added. Measure the difference in level of the mercury in the two tubes. This difference measures the pressure which the aqueous vapour is exerting.

From this experiment it appears that (the temperature remaining unchanged):—

(1) Water evaporates very rapidly in a vacuum.

(2) It is not possible for an unlimited amount of water to evaporate into a given space, but only a certain maximum amount. The space is then said to be *saturated*.

(3) When water is allowed to evaporate into a given space the aqueous vapour exerts a pressure, but this pressure cannot exceed a certain maximum.

What will be the effect of an increase in temperature upon the aqueous vapour pressure? Gently warm the upper end of the tube containing the water vapour with the hand, and then

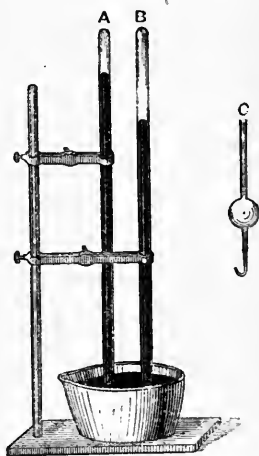


FIG. 50.

with a Bunsen burner. The mercury sinks still further. Hence :—

(4) As the temperature is gradually raised the pressure exerted by aqueous vapour is also increased. Can we measure the pressure corresponding to any particular temperature ?

**Dalton's Experiment for measuring the Pressure of Water Vapour.**—John Dalton was the first to carefully study

this subject, and his own account of his experiments may be quoted :

“I take a cylindrical glass tube open at both ends, and of two inches diameter and fourteen inches in length, to each end of which a cork is adapted, perforated in the middle so as to admit the barometer tube to be pushed through and to be held fast by them ; the upper cork is fixed two or three inches below the top of the tube, and is half cut away so as to admit water, &c., to pass by, its service being merely to keep the tube steady. Things being thus circumstanced, water of any temperature may be poured into the wide tube, and thus made to surround the upper part or vacuum of the barometer, and the effect of temperature in the production of vapour within can be observed from the depression of the mercurial columnn.”

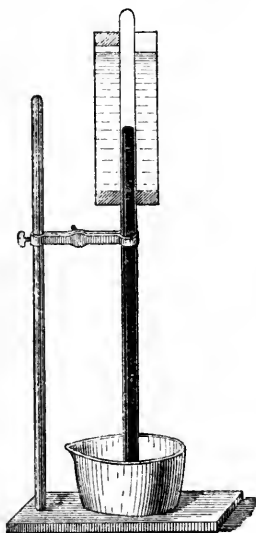


FIG. 51.

The experiments previously described show that water evaporates at ordinary temperatures below  $100^{\circ}$ . The pressure of such vapour formed at temperatures below  $100^{\circ}$  is often called *aqueous vapour pressure*, while the pressure exerted by the vapour arising from boiling water is commonly spoken of as *steam pressure*. The term *steam*, as commonly used, means either the vapour arising from boiling water or else aqueous vapour out of contact of water heated to temperatures above  $100^{\circ}$ .

The following table shows the pressure of aqueous vapour for

various temperatures, measured in millimetres of mercury, that is to say, by the number of millimetres of mercury through which the aqueous vapour at each temperature would depress a column of mercury.

#### PRESSURE OF AQUEOUS VAPOUR.

Temperature.	Pressure.	Temperature.	Pressure.
0° ... ..	4' + mm.	40° ... ..	54' + mm.
5 ... ..	6' +	60 ... ..	148' +
10 ... ..	9' +	80 ... ..	354' +
15 ... ..	12' +	90 ... ..	525' +
20 ... ..	17' +	100 ... ..	760' +

**Saturated and Unsaturated Aqueous Vapour.**—We have seen that if water be placed in a closed vessel the water evaporates very quickly if the vessel is vacuous, but that there is a certain maximum pressure corresponding to any given temperature which cannot be exceeded. If the vessel contains air it is found that the only difference is that the evaporation is slower: gradually the maximum pressure corresponding to the temperature is attained, and cannot be exceeded. In either case the air is said to be *saturated* with aqueous vapour.

Briefly, then, at any definite temperature the air is capable of taking up only a fixed amount of aqueous vapour.

Suppose the temperature of air which is saturated with aqueous vapour be reduced, what will be the consequence? At the reduced temperature the maximum vapour pressure is less than at the original temperature. Hence some of the aqueous vapour will have to be parted with. What can become of it? Will it not be condensed as water or dew again?

We are really well acquainted with many instances of this. In a hot room there may be a large amount of aqueous vapour, yet not sufficient for it to be saturated. Close to the windows on a cold night, however, the temperature of the air will be greatly reduced, so that there will be more than sufficient aqueous vapour present for it to be saturated, and as we know a heavy dew will be deposited upon the cold glass. So also the dew often found in the morning on the grass, and on the surface of leaf and flower and cobweb, comes from the cooling of the air around them to such a temperature that the invisible aqueous

vapour is in excess of that which can be contained in the air at this temperature, and it is therefore deposited.

If air containing aqueous vapour be slowly cooled, the temperature is sooner or later reached at which the air is saturated with vapour. The deposition of dew will begin at this temperature, and it is this temperature which is called the *dew-point*.

**Definition of the Dew-Point.**—*The dew-point is the temperature at which the aqueous vapour actually present in the air is sufficient to saturate it.*

**The Boiling of Water.**—In the above table the number 760 mm., representing the pressure of aqueous vapour at  $100^{\circ}$ , shows that at  $100^{\circ}$  the pressure of the vapour is sufficient to depress the mercury inside the tube to the level of the mercury outside. This means that the *pressure of aqueous vapour at  $100^{\circ}$  is equal to the normal atmospheric pressure.* Now it is exactly at this temperature that water evaporates in that rapid and violent way which we speak of as the *boiling* of water. This suggests that a liquid boils at a temperature such that the pressure of its vapour at that temperature is equal to the pressure of the atmosphere. How do experiment and experience bear this out?

If this be so, it follows that water should boil on a mountain at a lower temperature than at the sea level, for on the top of a mountain the atmospheric pressure is less than at sea level. This conclusion is borne out by Professor Tyndall's discovery that on the top of Mont Blanc water boils at  $85^{\circ}$ , or  $15^{\circ}$  below the temperature at which it boils at the sea level. Indeed for every 600 feet above sea level the boiling point of water is found to fall about one degree. But we do not need to ascend a mountain in order to lower the boiling point of water. The pressure of the air can be readily removed in other ways.

**EXPT. 2. Franklin's Experiment to show that Water will boil below its ordinary Boiling Point when the Pressure above it is lowered.**—Fit a strong round-bottomed flask with rubber stopper, glass tube, rubber tube and clip as shown in Fig. 52. Introduce some water and support the flask over wire gauze. Open the clip and boil the water vigorously till all the air is expelled, and its place taken by aqueous vapour. Then remove the flame, and close the clip. Allow the flask to cool until all boiling ceases. Then pour some cold water

over it. The water begins to boil. The reason of this is that a large portion of the aqueous vapour in the flask is condensed through the pouring on of the cold water. Consequently the pressure on the surface of the water is so far diminished that it is less than the maximum pressure of aqueous vapour at the temperature of the water, and the water therefore *boils*

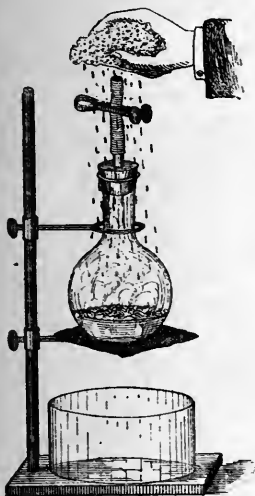


FIG. 52.

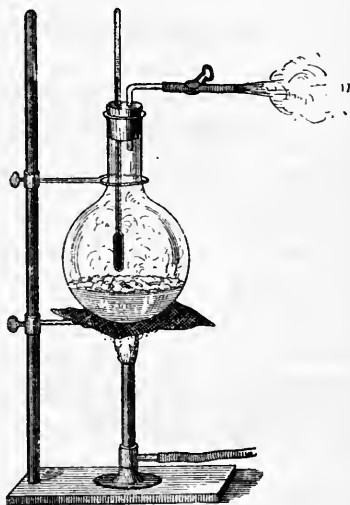


FIG. 53

### PROBLEM

Place a flask containing hot water under the air pump, or connect it with a Bunsen pump. Can you make the water boil?

Then again, supposing the pressure above the surface of water is increased, should not the boiling point be raised?

**EXPT. 3.<sup>1</sup> To show that the Boiling Point of Water can be raised above 100°.**—Pour some water into a strong flask, and close it with a rubber stopper, through which are passed a thermometer and a glass tube bent at right-angles (Fig. 53).

<sup>1</sup> Lecture Table Experiment.

Attach to the external end of the tube a piece of rubber tubing bearing a screw clip. Leaving the clip open, heat the flask on a flame until the water boils, and observe the temperature. Then close the clip until only a narrow exit is left for the steam. The result is that the pressure within the flask is gradually increased, and consequently the water boils at a higher temperature. As soon as the temperature is seen to rise to  $101^{\circ}$ , release the clip at once.

**Definition of the Boiling Point.**—*The boiling point of a liquid is that temperature at which the pressure of its vapour just overcomes the pressure of the atmosphere upon the surface of the liquid.*

**The Evaporation of other Liquids.**—We have considered the evaporation of water in some detail. Can other liquids evaporate at temperatures below their boiling points? can the statements made of aqueous vapour be applied to the vapours of other liquids than water?

**EXPT. 4. To compare the Evaporation of Liquids.**—Place drops of ether, chloroform, alcohol, benzene, water, and olive oil upon a slate. Observe that some disappear or are evaporated before others.

#### EXERCISE

Repeat Experiment 1, using (1) alcohol, (2) ether, (3) chloroform, in place of water.

The mercury sinks further when alcohol is used than with water, and still further in the case of ether, the temperature being the same in each case. Try also in each case the influence of an increased temperature.

It is evident then that the vapour pressure of alcohol or ether is greater than that of water at the same temperature. Moreover, in every case the vapour pressure increases as the temperature rises. It follows from this that with alcohol or ether the temperature will probably be sooner reached at which the vapour pressure is equal to the pressure of the atmosphere—in other words, the temperature at which the liquid will boil. If reference is made to the table of boiling points, p. 83, it is found in agreement with this conjecture that ether boils at  $35^{\circ}$ , and alcohol at  $78^{\circ}3$ .

**The Latent Heat of Vaporisation.**—We have found that while the vaporisation of a liquid is proceeding by boiling, although heat is being continuously added, nevertheless there is no simultaneous rise in temperature. This heat which apparently disappears but can be made sensible again is termed the *latent heat of vaporisation*.

In the process of boiling this heat is supplied from an external source.

Is heat also rendered latent when a liquid simply evaporates below the boiling point, when no heat is supplied from a flame or fire? If heat is absorbed in such a case, it will of necessity be obtained out of the liquid itself, or from surrounding objects, and consequently their temperatures will fall. We must test this by experiment and experience.

**EXPT. 5. Heat is rendered Latent in Ordinary Evaporation.**—Pour a little ether over the hand. A sensation of cold is felt.

Again recall the sensation of cold experienced when the body while wet after bathing is exposed to the air.

**EXPT. 6.** Take a small flask, fill it one-third full with ether. Place within it a test-tube containing about 1 c.c. of water. Then by means of a glass tube connected with an air-blast or bellows, send a rapid stream of air through the ether (Fig. 54). The ether is rapidly evaporated. Presently the water inside the tube is found frozen solid.

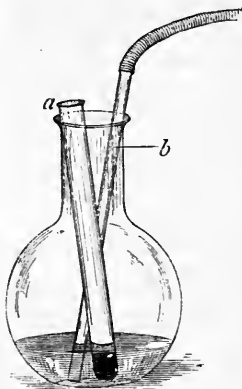


FIG. 54.

#### EXERCISE

Put some ether in a beaker. Wet the outside of the beaker with water. Stand it on a wooden block. Can you freeze the beaker to the block?

**EXPT. 7. The Cryophorus.**—This is an interesting piece of apparatus devised by the physicist Wollaston. It consists of

a closed glass tube, bent twice at right angles, with a bulb at each end, and contains some water and water vapour, but no air. Run all the water into the bulb B. Immerse the bulb A

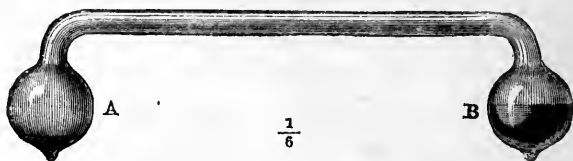


FIG. 55.

in melting ice. After a time the water in B is found to have become frozen. How do you account for this?

Several freezing machines have been devised which depend on the principle illustrated in this experiment.

**EXPT. 8. The Wet and Dry Bulb Thermometers.**—Take a thermometer. Wrap the bulb in cotton wool. Read its temperature. Pour a few drops of methylated spirit on the cotton wool. Wave the thermometer about in the air. Read the thermometer again. The temperature has fallen several degrees. Repeat, using water in place of alcohol.

It is a matter of common knowledge that the air is much more moist on some days than on others. Sometimes the air appears to be laden with moisture, and everything out of doors is dank and wet. On such a day it would be useless to hang clothes out to dry, for the air is saturated or contains as much aqueous vapour as is possible at that temperature, and if a wet thermometer bulb were moved about in such air no fall of temperature would occur. It is evident therefore that we can judge of the un-saturation of the atmosphere by observing whether a moist thermometer sinks below the actual temperature of the air or not. The *wet and dry bulb thermometer* (Fig. 56) is an instrument which enables us to do this. One thermometer serves to record the actual temperature of the air. The bulb of the other is surrounded by a piece of muslin one end of which dips into a small vessel containing water, and by these means the bulb is kept moist. On a dry day the water on the wet bulb evaporates, and a lower temperature is recorded than by the dry thermometer. Moreover, the magnitude of the difference between the



readings of the two thermometers gives a measure of the degree of un-saturation of the atmosphere, for the drier the air, the more rapid is the evaporation and the greater the lowering of temper-

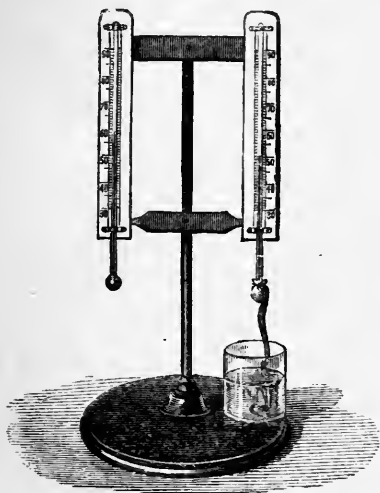


FIG. 56.

ature. On a thoroughly damp day the two thermometers record the same temperature

#### PROBLEM

Hang up a piece of salt seaweed. Compare its condition, day by day, with the readings of the wet and dry bulb thermometer.

## CHAPTER XIII

### CALORIMETRY

**The Unit of Heat.**—We have already seen that when a solid is melted, or when a liquid is vaporised, heat becomes latent (pp. 78, 91), but we cannot endeavour to measure the *quantity* of heat which becomes latent until we have fixed upon some standard measure, that is upon a definite *unit* of heat.

The effects of heat are numerous, and any one of them which can be accurately measured might be used as a unit of heat ; for instance, we might select as our unit of heat the quantity necessary to produce a definite increase of volume in unit mass of some standard substance. As a matter of fact the effect of heat selected is that of increase of temperature, and the unit of heat is the *quantity of heat required to raise the temperature of 1 gram of water from  $0^{\circ}$  C. to  $1^{\circ}$  C.* ; this unit is usually called *one calorie*. Hence, if 50 grams of water are to be heated from  $0^{\circ}$  to  $1^{\circ}$  C., 50 units of heat will be required.

Pending further inquiry we shall assume that the quantities of heat necessary to raise the temperature of 1 gram of water one degree are equal wherever that degree may be on the scale of temperature ; for instance, the quantity of heat necessary to raise the temperature of 1 gram of water from  $80^{\circ}$  to  $81^{\circ}$  may be taken as equal to that required to raise the temperature of 1 gram of water from  $0^{\circ}$  to  $1^{\circ}$ . It follows from this that if the temperature of 1 gram of water be raised from  $0^{\circ}$  to  $80^{\circ}$ , 80 calories are required, and if the temperature of 20 grams of

water be raised from  $30^{\circ}$  to  $55^{\circ}$ ,  $25 \times 20$  or 500 calories will be required.

It must be assumed also that when a given weight of water or indeed of any substance is cooled through any range of temperature it gives up exactly the same quantity of heat that is required to raise its temperature through the same range.

Let us next consider what occurs when we mix together two quantities of water at different temperatures.

**EXPT. I. The Effect of mixing two Quantities of Water at Different Temperatures.**—Support a small cylindrical vessel *a*, made of sheet brass<sup>1</sup> on corks inside a larger zinc vessel *b* (Fig. 57). The inner vessel is called a "calorimeter" or heat-measurer. Measure into it 50 c.c. of water. Place a thermometer inside it. Weigh a small flask. Measure 50 c.c. of water into it. Place another thermometer inside it,

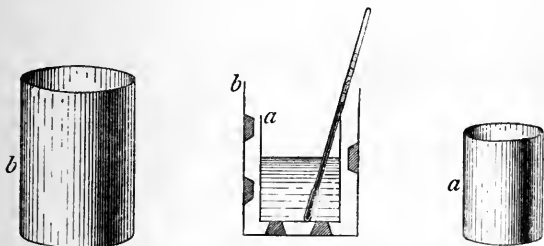


FIG. 57.

and heat the flask slowly over a small flame. Meantime stir the water in the calorimeter and read its temperature. As soon as the water in the flask is at about  $50^{\circ}$  remove the flame, stir carefully with the thermometer, quickly read its temperature, take out the thermometer, and pour the water into the calorimeter. Stir the mixture in the calorimeter, and read the highest temperature attained to one-tenth of a degree if possible. Weigh the flask and so determine the quantity of water left

<sup>1</sup> A small glass beaker may be used if a brass vessel is not available.

in the flask. The following numbers were obtained in an experiment :—

Weight of warm water measured out ... ..	50	gram.
"          "          left in flask ... ..	0·5	"
"          "          used... ..	49·5	"
Temperature of warm water ... ..	52°·1	C.
Weight of water originally in calorimeter ... ..	50	gram.
Temperature of water in calorimeter ... ..	19°·1	"
Temperature of mixture observed ... ..	35°·1	"

It is evident that the water in the calorimeter gained heat while heat was lost by the warm water added to it. If there was no loss of heat and no interchange of heat except between the two quantities of water we can readily calculate what the common and final temperature should be, and this should agree with the final temperature observed, viz., 35°·1.

Probably the simplest and most convenient way is to calculate the quantities of heat which both quantities of water contained *above* that they would have possessed were their temperatures 0°. These were :—

$50 \times 19\cdot1$  or 955 calories in the case of the water in the calorimeter.  
 $49\cdot5 \times 52\cdot1$  or 2578·95 calories in the case of the warm water.

Hence,

$955 + 2579$  or 3534 calories of heat were absorbed by  $50 + 49\cdot5$  or 99·5 grams of water.

Hence, the common temperature should be  $\frac{3534}{99\cdot5}$  or 35°·5,  
 whereas it was actually found to be 35°·1.

The difference between the observed and calculated values is probably due to a slight loss of heat from the water; for instance, a small amount of the heat in the warm water goes to heat not only the water in the calorimeter, but also the calorimeter itself and the thermometer. (*See* Expt. 4, p. 99.)

## SPECIFIC HEAT.

When equal quantities of heat are communicated to equal weights of water, the temperature of each is raised by the same amount. But if equal weights of *different* substances are heated until they have acquired the same temperature, will they be found to have absorbed the same quantities of heat? In other words, have different substances the same *capacity* for heat?

**EXPT. 2. To find whether Different Metals have the same Capacity for Heat.**—Take cylinders of lead, iron, and copper, of the same weight, say fifty grams in weight each. Suspend them by threads in boiling water for five minutes. Take three calorimeters and measure 50 c.c. of water at the ordinary temperature into each. Place a thermometer in each. Then remove the cylinders at the same time from the boiling water, and place one in each calorimeter. Stir carefully with the thermometer, and read in each case the highest temperature attained. Have the metals heated the water to different temperatures? Which has raised the temperature of the water the most?

It will be found from this experiment that if equal quantities of heat are given to equal weights of *different* substances, the temperatures of each will be raised by different amounts.

**Definition of Specific Heat.**—*The quantity of heat required to raise the temperature of 1 gram of a substance from  $0^{\circ}$  C. to  $1^{\circ}$  C. is called its specific heat at  $0^{\circ}$ . Since, however, very nearly the same quantity of heat is required to raise the temperature of any substance one degree wherever that degree may be on the scale of temperature, we may state as approximately true that the specific heat of any substance is the quantity*

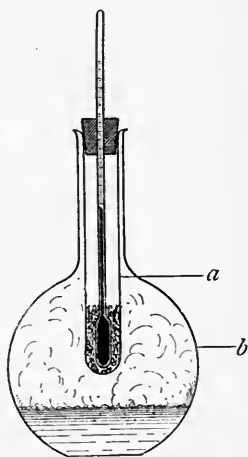


FIG. 58.

of heat necessary to raise the temperature of one gram of it  $1^{\circ}\text{C}$ .

EXPT. 3. To find the Specific Heat of a Metal.<sup>1</sup>—Measure into a thin brass vessel<sup>2</sup> 100 c.c. of water and place a thermometer in it. Weigh into a test-tube *a* (Fig. 58) about 50 grams of the metal. Close it with a stopper, through which is passed a thermometer, the bulb of which must be surrounded by the metal. Heat the metal as nearly as possible to  $100^{\circ}$  by suspending the test-tube for some time in a flask *b* containing boiling water. Now transfer the metal as quickly as possible into the water in the beaker, stir with the thermometer, and observe the highest temperature reached. Thus in an experiment the following data were obtained :—

Weight of metal	...	...	...	...	...	...	...	50 gm.
„ water	...	...	...	...	...	...	...	100 gm.
Temperature of metal before experiment	...	...	...	...	...	...	...	$100^{\circ}\text{C}$ .
„ cold water before experiment	...	...	...	...	...	...	...	$12^{\circ}\text{C}$ .
„ water after experiment	...	...	...	...	...	...	...	$16^{\circ}\text{C}$ .

Assume *calories lost by metal = calories gained by water* ; then,

Metal fell in temperature	...	...	...	...	...	...	...	$100 - 16$ or $84$ degrees.
Water increased in temperature	...	...	...	...	...	...	...	$16 - 12$ or $4$ degrees.
Water gained	...	...	...	...	...	...	...	$100 \times 4$ or $400$ calories.
i.e., metal lost	...	...	...	...	...	...	...	$400$ calories.

If 50 gm. metal cooling  $84^{\circ}$  lost 400 calories,

$$1 \text{ gm. } ,, ,, 1^{\circ} \text{ lost } \frac{400}{50} \times \frac{1}{84} \text{ or } \frac{8}{84} \text{ calories.}$$

Hence,

$$\text{Specific Heat of metal} = .09 +$$

The student should observe that (1) the metal must be thoroughly heated, (2) as the difference of temperature to be measured is very small, temperatures should be read to the nearest tenth of a degree.<sup>3</sup>

<sup>1</sup> e.g., brass chain, lead shot, iron nails, granulated zinc, copper turnings.

<sup>2</sup> A thin glass beaker may be used.

<sup>3</sup> *Note for Teacher*.—The method described above is only moderately accurate but it is instructive, and does not take too long. The student realises at least what is meant by “specific heat,” which is the important thing. It may be noted that the “water equivalent” is neglected, and that it is assumed that 100 c.c. = 100 grams.

**Water Equivalent of a Calorimeter.**<sup>1</sup>—If warm water is added to a calorimeter, containing a thermometer and stirrer, which is at the temperature of the air, the calorimeter, stirrer and thermometer are heated, and on the other hand there is a loss of heat from the water, which must be allowed for in accurate determinations of specific heat. Now there must be a certain quantity of water which will absorb just as much heat as the calorimeter, stirrer, and thermometer absorb in being heated through any range of temperature. This equivalent amount of water is called the *water equivalent of the calorimeter*, and must be allowed for in accurate determinations of specific heat.

**EXPT. 4.**<sup>2</sup> **To show the Absorption of Heat by a Calorimeter.**—Observe the temperature of the room. Place a small brass stirrer in a brass calorimeter. Weigh. Support the calorimeter as before on cork in the outer zinc vessel. Heat some water in a flask containing a second stirrer and a thermometer to about 35°. As soon as this temperature is reached, remove the flame, stir well, read the temperature, and quickly pour the water into the calorimeter until it is about one-third full. Then quickly put the thermometer in the calorimeter, stir well, watching closely the thermometer. The temperature falls rapidly through a small range owing to the absorption of heat by the calorimeter and stirrer. Read the temperature. From this point the temperature continues to fall slowly owing to the cooling of the apparatus.

For a fuller treatment of this subject the student should refer to books on Physics.

TABLE OF SPECIFIC HEATS.

SOLIDS.

Magnesium ... ..	0·250	Zinc ... ..	0·096
Aluminium ... ..	0·214	Copper ... ..	0·095
Glass ... ..	0·188	Silver ... ..	0·057
Sulphur... ..	0·177	Tin ... ..	0·056
Diamond ... ..	0·147	Gold ... ..	0·032
Iron ... ..	0·114	Lead ... ..	0·031

<sup>1</sup> Omit on first studying the subject.

<sup>2</sup> Omit on first studying the subject.

## LIQUIDS.

Water ... ..	1.000	Oil of turpentine...	0.462
Alcohol ... ..	0.615	Mercury ... ..	0.033
Ether ... ..	0.503		

It is seen at once from these numbers that the specific heat of water very greatly exceeds that of most other substances.

## LATENT HEAT OF LIQUEFACTION.

We have already found that heat becomes latent when any solid substance is melted, and we are now in a position to measure the latent heat of liquefaction of ice. But first we must define exactly what it is we wish to measure.

**Definition of Latent Heat of Liquefaction.**—*The latent heat of liquefaction of any substance is the number of calories of heat absorbed by one gram of it in passing from the solid to the liquid state without rise of temperature.*

N.B.—This is equal to the number of calories of heat evolved when a gram of the liquid substance is reconverted to the solid state without fall of temperature.

**EXPT. 5. To measure the Latent Heat of Liquefaction of Ice.**—Weigh a small sheet-brass calorimeter. Add 100 c.c. of water, and weigh again. Heat it carefully to about 20°. Then weigh out roughly on filter-paper about 10 grams of ice. Place the calorimeter on cork supports within a zinc can, and carefully read the temperature of the water in it. Dry the ice with blotting paper. Quickly introduce it into the calorimeter, stir thoroughly and observe the lowest temperature attained. Finally, remove the thermometer, allowing the water to drain off into the calorimeter, and weigh again. The increase in weight gives the weight of ice added. In an experiment the following numbers were obtained :—

Weight of calorimeter ... ..	21.4	gram.
„ calorimeter and water ... ..	119.2	„
„ calorimeter, water and ice ... ..	128.5	„
Initial temperature of water ... ..	21.0	C.
Final „ „ ... ..	12.5	C.
Water equivalent of calorimeter ( previously determined) ... ..	1.84	gram.



Thus,  $119.2 - 21.4$  or  $97.8$  grams of water were cooled from  $21^{\circ}$  to  $12^{\circ}.5$ .  
 $128.5 - 119.2$  or  $9.3$  grams of ice were added.

This ice *was first melted, and secondly heated to  $12^{\circ}.5$ .*

Assume

heat gained by water = heat lost by ice.

Then,

$$\begin{aligned}\text{heat lost by water} &= 97.8 \times (21 - 12.5) \\ &= 831.3 \text{ calories.}\end{aligned}$$

therefore,

$$\text{heat gained by } 9.3 \text{ grams ice} = 831.3 \text{ calories.}$$

$$\begin{aligned}\text{But, No. of calories gained by } 9.3 \text{ grm. of melted} & \\ \text{ice at } 0^{\circ}, \text{ due to } 12^{\circ}.5 \text{ rise of temperature} & \} = 9.3 \times 12.5 \\ & = 116.25\end{aligned}$$

$$\begin{aligned}\text{Hence, the number of calories absorbed in} & \\ \text{melting } 9.3 \text{ grams of ice} \dots \dots \dots & \} = 831.3 - 116.25 \\ & = 715.05\end{aligned}$$

Therefore,

$$\begin{aligned}\text{number of calories absorbed in melting} & \\ \text{1 gram of ice} \dots \dots \dots & \} = 76.9\end{aligned}$$

The value obtained here for the latent heat of liquefaction of ice must be slightly inaccurate, owing to a small loss of heat from the apparatus, and to the ice being slightly wet on the surface. The true value lies between 79 and 80, but in round numbers it may conveniently be taken as 80. We have arrived then at the result that in the change of 1 gram of ice at  $0^{\circ}$  into 1 gram of water at  $0^{\circ}$  no less than 80 calories of heat are rendered latent; that is to say, as much heat as could heat 80 grams of water from  $0^{\circ}$  to  $1^{\circ}$ , or heat 1 gram of water from  $0^{\circ}$  to  $80^{\circ}$ .

It is interesting to observe the values of the latent heats of liquefaction of other bodies.

TABLE OF LATENT HEATS OF LIQUEFACTION.

Ice	...	...	...	80		Tin	...	...	...	14 +
Sodium nitrate	...	...	...	63		Iodine	...	...	...	11 +
Nitre	...	...	...	47 +		Sulphur	...	...	...	9 +
Zinc	...	...	...	28 +		Lead	...	...	...	5 +
Silver	...	...	...	21 +		Phosphorus	...	...	...	5 +

It is very noticeable that the latent heat of ice is very great as compared with those of other substances, and this is a matter of

the greatest importance. If snow and ice were immediately converted into water at  $0^{\circ}\text{C}$ . with little or no absorption of heat, the sudden melting of the snow on the hills and mountains would expose the inhabitants of the valleys below them to the most terrible floods and inundations.

### LATENT HEAT OF VAPORISATION.

Just as heat disappears whenever a solid is liquefied, so also we have already seen that when a liquid is vaporised heat becomes latent.

**Definition of Latent Heat of Vaporisation.**—*The latent heat of vaporisation of any liquid at a given temperature is the number of calories of heat absorbed by 1 gram of it in passing from the liquid to the gaseous state without rise of temperature.*

When a vapour is condensed again into a liquid the heat which was insensible or latent becomes sensible again, and the number of calories evolved is equal to the number previously absorbed.

**EXPT. 6. To Measure the Latent Heat of Vaporisation of Water.**<sup>1</sup>—Place some water in a 500 c.c. flask. Close it with a stopper, through

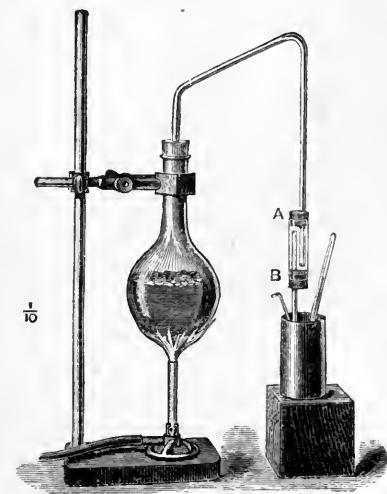


FIG. 59.

which are passed a thistle-funnel (its lower end dipping below the surface of the water, not shown in Fig. 59) and a glass tube bent as shown. Proceed to heat the water. Also take a wide glass tube A B, closed by stoppers through the lower of which a straight glass tube is passed. Connect this apparatus,

<sup>1</sup> Lecture Table Experiment.

which is to serve as a trap, with the flask by means of the bent glass tube. The trap serves to catch the condensed water which is carried over by the steam. When the steam issues freely from the end of the tube, measure 50 c.c. of ice-cold water into a weighed brass calorimeter or beaker, and place it so that the end of the tube reaches nearly to the bottom of the calorimeter.

The steam is condensed with a loud noise by the cold water and the temperature rapidly rises, till finally the water in the beaker itself is raised to near the boiling point. Then remove the flame and also the beaker, and observe that the volume of the water in the beaker is only increased by about one-fifth. If the increase in weight of the beaker is 10 grams, this shows that 10 grams of steam at  $100^{\circ}$ , in changing into 10 grams of water at  $100^{\circ}$ , set free sufficient heat to raise the temperature of 50 grams of water from  $0^{\circ}$  to  $100^{\circ}$  C. What value does this result give for the latent heat of vaporisation of water?

This lecture table experiment shows that the latent heat of vaporisation of water has a high value, but no really accurate result can be obtained with this apparatus, without taking many precautions.

The true value of the latent heat of vaporisation of water at  $100^{\circ}$  is found to be 537; that is to say, in converting 1 gram of water at  $100^{\circ}$  into steam at  $100^{\circ}$ , no less than 537 units of heat are rendered latent.

TABLE OF LATENT HEATS OF VAPORISATION

Water	...	...	537	Oil of turpentine	...	74
Alcohol	...	...	202 +	Bromine	...	45 +
Ether	...	...	105 +	Iodine	...	24
Acetic acid	...	...	102			

It is seen from this table that the latent heat of vaporisation of water is exceptionally high and greatly exceeds that of almost every other liquid, just as the value of its specific heat and that of its latent heat of liquefaction are higher than those of other liquids.

In view of the fact that water exists in vast quantities in the liquid form on the surface of the globe, that its vapour is always

present in the air, and that ice and snow are forms in which it is often found, there can be no doubt that the above exceptional properties of water are of the highest importance in maintaining a fairly uniform temperature on the surface of the earth.

## EXAMPLES XI

### Specific Heat

- Find the Specific Heat of a silver watch chain having given  

Weight of chain	...	...	...	...	...	90 grams.
Temperature of hot chain	...	...	...	...	...	99° C.
Volume of cold water	...	...	...	...	...	100 c.c.
Temperature of cold water	...	...	...	...	...	15° C.
Temperature of mixture	...	...	...	...	...	19°·1 C.

2. If 2 lbs. of mercury at 100° C. are mixed with 1 lb. of water at 7° C. and the resulting temperature is 13° C., find the Specific Heat of mercury.

3. 10 grams of a certain substance are heated to 100° C. and placed in 75 grams of water at 4° C. The final temperature of the water is 10° C. Find the Specific Heat of the substance.

4. Find the final temperature of a mixture of 100 grams alcohol at 10° C. and 100 c.c. water at 0° C., assuming that the S. H. of alcohol is ·6, and that the only heat change is an interchange between the alcohol and the water.

5. Find the final temperature of 100 grams of water originally at 0° C., after dropping into it 10 grams of copper nails at 100° C. (S. H. of copper = ·095.)

6. Find the resulting temperature when 10 grams of water at 100° C. are poured into a copper vessel at 0° C. weighing 100 grams. (S. H. of copper = ·095.)

7. What would be the rise of temperature on dropping 100 grams of lead shot at 100° C. into 100 grams of water at 0° C.? (S. H. of lead = ·03.)

8. How many grams of silver at 90° if dipped into 100 grams alcohol at 10° would raise its temperature to 20° C.? (S. H. of silver = ·057; alcohol = ·615.)

9. What weight of mercury at 100° if mixed with 100 grams of water at 0° would raise its temperature to 10° C.? (S. H. of mercury =  $\frac{1}{35}$ .)

## CHAPTER XIV

### SOLUTION

**THE SOLVENT ACTION OF WATER.**—What occurs when a solid substance is placed in water? Place a lump of sugar in a small beaker of warm water. The sugar crumbles down, melts and disappears in the water. But the sugar is not lost, for the water tastes sweet.

This sort of melting and disappearing in the presence of water is called *solution*, and the water is said to have a *solvent action* upon the substance dissolved. All substances, however, do not dissolve so readily in water.

**EXPT. 1. To find whether Plaster of Paris is soluble in Water.**—Shake some powdered plaster of Paris in a small flask with water. The water becomes milky owing to powder *in suspension*. The plaster seems insoluble.

But does the evidence justify this conclusion? We know that the plaster *is not all dissolved*. Is it possible that some of it is dissolved? To settle this we must separate the liquid from the floating powder. This we can do by letting it stand, but that takes time. Can we not pour the milky liquid on some porous material such as blotting-paper, which will allow the liquid part to run through while the floating solid particles are caught and retained on the paper? Try.

**EXPT. 2. To Filter the Milky Liquid containing Plaster of Paris.**—Take a circular piece of a kind of porous paper called filter-paper. Fold it once across a diameter. Then fold it again in two. Open the folded paper so that it forms a cone as shown in Fig. 60. Place it in a dry funnel, the cone

of which is larger than the paper cone. It springs up if the paper does not fit the glass closely all round ; alter the folding till it fits well. Now moisten the paper with water to keep it in place. Place the funnel in the mouth of a flask, or in the ring of a retort-stand above a clean beaker.

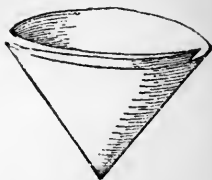


FIG. 60.

Take the flask containing the water which has been shaken up with the plaster of Paris. Pour the contents down a glass rod into the funnel (*see* Fig. 61), and take care that the liquid does not

rise above the top of the paper. The plaster is retained on the paper, while a clear liquid runs through ; this latter is called the *filtrate*. If the first few drops are not quite clear, pour them on the filter again.

If the clear filtrate is evaporated the water will be driven off. Solid plaster of Paris does not evaporate into the air. It seems probable, then, that if the clear filtrate be evaporated, solid plaster of Paris will be left behind, if the water has really dissolved any of it.



FIG. 61.

**EXPT. 3.** To find whether the clear Filtrate contains dissolved Plaster of Paris.—Place a drop of the filtrate on

platinum foil. Hold the foil by a pair of tongs over a flame. A white stain is left. It is clear, then, that plaster of Paris is *slightly soluble* in water. We may evaporate a large quantity of the filtrate as follows.

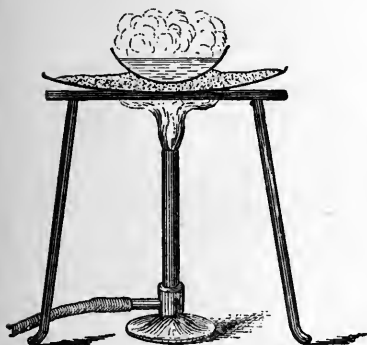


FIG. 62.

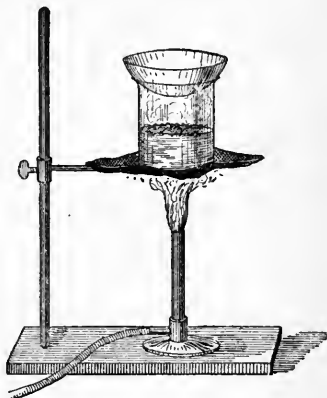


FIG. 63.

EXPT. 4. Pour some of the filtrate into a glass basin. Place the basin on a sand-bath (*see* Fig. 62), and heat it so that the liquid is evaporated without spurting. The volume of liquid gradually lessens, and a white deposit crusts the basin.

A still surer method for avoiding loss by spurting is by evaporating the liquid on a "*water-bath*." The basin is in this case put on a beaker in which water is being boiled, as shown in Fig. 63.

#### PROBLEMS

Examine the solvent action of water upon nitre, salt, green vitriol, garden soil, gravel, chalk.

**SOLIDS IN NATURAL WATERS.**—Do your results not bring out the fact that several substances which occur in nature are soluble in water? Then will not water which has flowed over or through the crust of the earth be found to have dissolved something? We can readily find out whether this is so.

EXPT. 5. To find whether Spring Water contains Dissolved Solids.—Weigh a glass evaporating basin. Measure 100 c.c. of clear spring water into it. Evaporate the water by

placing the basin on a sand-bath or water-bath. Weigh again after the disappearance of the water.

A solid residue is seen to be left, and a small increase in weight will be observed.

### PROBLEMS

Find whether rain water, river water, sea water,<sup>1</sup> do or do not leave a solid residue on evaporation.

You have probably found that while the rain water leaves little or no residue the sea water leaves very much more than the spring or river water.

Then rain water—*i.e.*, water which has not been in contact with the surface of the earth—is almost entirely free from dissolved solid impurity. Spring and river waters, on the other hand, are waters which have trickled through or over various strata of the earth's crust, so that it is probable that much of the solid matter they contain has been dissolved out of the rocks and soil over which they have passed. This will account, too, for sea water containing so much dissolved matter. For the rivers flow down into the sea, carrying with them that which they have dissolved out of the earth, and as the water of the sea evaporates the solid material remains behind, just as in the above experiment.

It is found that *river water* may contain 7 to 55 grams of solid matter per 100,000 c.c. of water, while *spring water* usually contains from 3 to 150 grams. In contrast to these the water of the mid-Atlantic Ocean contains about 3,400 grams per 100,000 c.c.

### SOLIDS IN GRAMS PER 100,000 C.C. OF NATURAL WATERS

Rain ... average	3.5 grams.	Lakes:—	
Rivers:—		Loch Katrine ...	28 grams.
Neva ... ..	5.5 „	Thirlmere ...	27 „
Irwell near source	8 „	Mineral Waters:—	
„ below Man-		Vichy ... ..	503 „
chester ...	56 „	Eau de Seltz ...	407 „
Rhine ... ..	16–32 „	Sea Water:—	
Thames, at London,	40–45 „	Black Sea ...	1,770 „
Jordan ... ..	105 „	Indian Ocean ...	3,400 „
		North Atlantic	
		Ocean ... ..	3,849 „
		Dead Sea ...	17,000 „

<sup>1</sup> Sea water is supplied by most chemical firms.



**HARD AND SOFT WATERS.**—The differences which are found between the amounts of solid matter contained in various natural waters are notable, and of much importance.

Waters which contain much dissolved matter usually yield a lather with difficulty with soap, so that much soap is used up before its cleansing action can be utilised.

**EXPT. 6. To make a Soap Solution.**<sup>1</sup>—Cut a piece of Castille soap so that it weighs 10 grams. Cut it into thin slices, and place them in a large flask. Add 400 c.c. methylated spirits. Warm gently and shake until the soap is dissolved. Pour the solution into a litre flask. Wash out the other flask with a little alcohol. Add distilled water so as to make the volume up to 1 litre. Store the solution in a Winchester pint bottle.

**EXPT. 7. To compare the Hardness of Different Waters.**—Fill a burette with the soap solution. Take a 6-oz. glass-stoppered bottle. Measure into it 100 c.c. of the laboratory tap water. Read the burette. Run into the bottle 1 c.c. of soap solution. Close with the stopper and shake well. Place the bottle on its side on the bench. Is there a permanent lather? No! Then, add another 1 c.c. and shake, and proceed in this way until a lather remains right across the surface of the water for three minutes. Read the volume of soap solution added.

Find in the same way what volumes of the same soap solution must be added in order to produce a permanent lather with different kinds of water—*e.g.*, (*a*) rain water, (*b*) a river water, (*c*) a spring water.

You probably find that waters, which you know contain much solid matter, lather with difficulty. Such waters are called *hard* and are ill adapted for washing purposes. Waters which contain little dissolved matter, on the other hand, lather readily. They are termed *soft*.

**EXPT. 8. A Hard Water when boiled deposits a Solid Crust.**—Take some water which you know to be very hard. Half fill a perfectly clean 8-oz. flask with the water. Boil steadily for some minutes. Then examine the inner surface of the flask. A thin film of solid matter has been deposited wherever the water was in contact with the glass. Examine

<sup>1</sup> Lecture Table Experiment.

the inside of a kettle in which hard water has frequently been boiled, and note the fur on the sides.

Thus it is clear that a boiler in which immense quantities of water are heated will in the course of time become coated with a thick solid crust, if the water is hard. This leads sometimes to dangerous accidents because the iron, if protected by the incrustation from the water, becomes greatly overheated. Hence, a water which contains a considerable quantity of solid matter in solution is generally ill suited for use in boilers.

It may be added that soft waters are much better adapted than hard waters for cooking purposes, whether for making tea and bread, or boiling vegetables and meat. On the other hand, it is found that soft waters have the disadvantage of attacking metals, especially lead, much more than hard waters do.

**Distilled Water.**—We have seen that a solid residue is left when a water containing solid matter is evaporated. If, as we have really assumed, the whole of the dissolved matter is left behind, the water which evaporates must be free from such impurity.

**EXPT. 9. Pure Water can be obtained by Distillation.**—(1) Distil about 200 c.c. of sea water from a flask, as in Expt. 4, p. 81. Collect 50 c.c. of the distillate, and show that on evaporation it leaves no residue.

(2) Colour about 200 c.c. of water with a little ink, and distil the liquid. The distillate is quite colourless.

We have, therefore, in the process of distillation a means of obtaining pure water from such water as sea water, which contains much dissolved solid matter, and is unfit for drinking purposes.

**Rain Water is Distilled Water.**—Distillation is always occurring in nature. From the wide expanse of the waters of the ocean, evaporation is continually taking place. Leaving the dissolved matter behind, the water vapour ascends to descend again in the form of pure rain water upon the surface of the earth.

**THE SOLUBILITIES OF SOLIDS IN WATER.**—Water, we have seen, can dissolve many different substances. Are all substances, however, soluble to the same extent in water?

Sugar is readily soluble: plaster of Paris only slightly so. But we must make actual measurements with different substances.

**EXPT. 10. To find what Weight of Salt can be dissolved by Water.**—Half fill a 4 oz. flask with pure water at the ordinary temperature. Add some common salt, and shake well. If it is entirely dissolved, add more until some remains undissolved. Note that it is impossible to dissolve an unlimited amount in the water. At the temperature of the experiment the water will only dissolve a certain amount of the salt. The solution is then said to be *saturated*. Allow it to stand. The undissolved salt quickly settles, leaving the solution above it clear. Weigh an evaporating basin or clock-glass. Carefully suck up 20 c.c. of the clear solution into a pipette, and transfer to the basin. Evaporate the solution to dryness on a water-bath. Weigh. Calculate the weight of salt dissolved in 100 c.c. of the solution. In an experiment,

20 c.c. solution contained 5.91 grm. salt.

Hence, 100 c.c. „ „ 29.55 grm. salt.

**Definition of Solubility.**—*The solubility of a substance in water at any given temperature is measured by the weight of the substance which can be dissolved in 100 grams of water at that temperature.*

Notice, however, that in the last experiment we measured the weight of salt dissolved in 100 c.c. of the solution, not in 100 grams of water. In order to find the solubility we must begin by *weighing* the solution taken. Thus, it was found that 20 c.c. salt solution weighed 22.21 grams, and contained 5.91 grams salt.

Hence,

22.21 – 5.91 or 16.3 grm. water dissolved 5.91 grm. salt.

Therefore, 100 grm. water dissolved  $\frac{5.91}{16.3} \times 100$  or 36.2 grm. salt.

### PROBLEMS

Find what weight of each of the following substances can be dissolved by 100 grams of water, saturated at the ordinary temperature:—powdered nitre, chalk, sand, garden soil, powdered gypsum.

Do not your results bring out the important fact that substances of common occurrence differ greatly as regards their solubility in water?

**The Influence of Temperature upon Solubility.**—Thus far we have only examined the solubilities of substances in water at the ordinary temperature. But is it not possible that at higher temperatures the solubility of a substance will be different from what it is at a low temperature? We must, then, study the influence of temperature.

**EXPT. II. Nitre is more soluble in Hot than in Cold Water.**—Half fill a test-tube with water. Add sufficient powdered nitre to saturate it at the ordinary temperature. Then heat the test-tube in a flame. The undissolved nitre dissolves. Add more nitre. It can also be dissolved. Hold the test-tube in a stream of cold water from a tap, and at the same time scratch the inner surface of the tube with a glass rod. As the temperature of the solution falls, some of the dissolved nitre separates out in crystals.

Then, there can be no doubt that nitre is more soluble in hot water than in cold water.

### PROBLEMS

Find whether blue vitriol, alum, bichromate of potash, chlorate of potash, common salt, powdered gypsum, are more soluble in cold or in hot water.

**Solubility Curves.**<sup>1</sup>—The results of careful experiments on the amounts of nitre which can be dissolved in 100 grams of water at different temperatures can be best expressed by a graphical method.

Draw two lines A B and A C at right angles (Fig. 64) Divide A B into ten equal parts and mark them  $10^{\circ}$ ,  $20^{\circ}$ ,  $30^{\circ}$ , etc., and divide A C into ten equal parts and mark them 10 gram, 20 gram, etc. Then the mark X indicates the result of an experiment which shows that 100 grams water will dissolve 20 grams of nitre at  $10^{\circ}$ . So the mark Y indicates that 100 grams of water will dissolve 32 grams of nitre at  $20^{\circ}$ . A curve K N is obtained if these marks and those representing other

<sup>1</sup> Omit on a first study.

experiments be joined by a line. This line, called the *solubility curve* of nitre, enables us to read off the amount of nitre which 100 grams of water will dissolve at any temperature between  $0^{\circ}$  and  $100^{\circ}$ .

Several such curves can be shown in one diagram. In Fig. 65 are given the solubility curves of nitre, chlorate of potash and common salt. It is seen, for instance, that below  $25^{\circ}$ ,

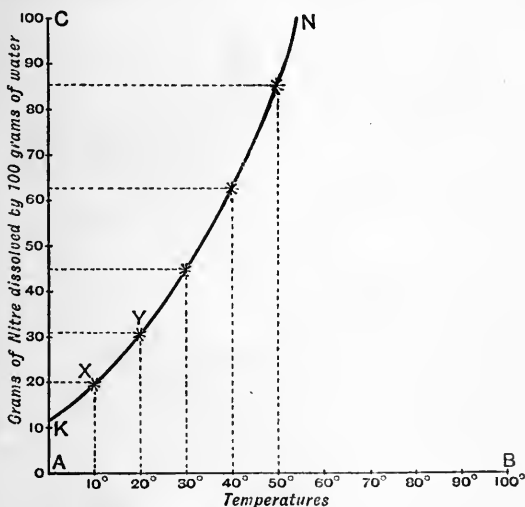


FIG. 64.

common salt is more soluble in water than nitre is ; above  $25^{\circ}$ , however, the reverse is the case.

**Solutions of known Strength.**—It is sometimes necessary to prepare a solution of known strength, that is, one containing a known quantity of a substance in a given volume. The method of doing this may be illustrated by an example.

**EXPT. 12. To prepare a Solution containing 50 Grams of Nitre per Litre.**—Your object will be to make the solution of the exact strength required without needless loss of time. Grind the nitre to powder. This will help it to dissolve.

Heat a little gently in a dry test-tube. If any moisture condenses on the side of the glass, this shows that the nitre is damp. Place the nitre in an evaporating basin over a small flame. Stir with a glass rod until dry. If the nitre is weighed on the balance pan or even on a sheet of paper some may be lost in transferring it to the vessel in which the solution is made. Counterpoise an empty flask. Weigh into it 50 grams of nitre. Add about 200 c.c. of hot water. You will naturally have saved time by putting water on to boil whilst weighing ; besides, any attempt to boil cold water in a flask at the bottom of which

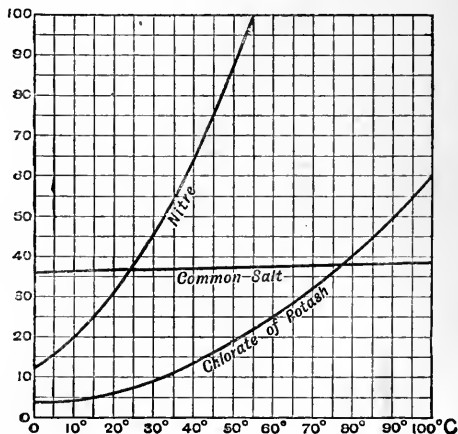


FIG. 65.

there is a thick layer of undissolved solid is liable to crack the glass. [Why?] Shake and boil until all the nitre is dissolved. The liquid is now hot and will contract on cooling. Run water over the outside of the flask to cool it. Pour the solution into a litre measuring flask. Rinse the flask three times. Dilute to the mark with cold water.

Is there any objection to filtering a solution before it is diluted? Could you get over the objection?

If a solution has to be both cooled and diluted, which should be done first, and why?

**The Solvent Action of other Liquids besides Water.**

—We have up to the present studied the solvent action of water alone upon solids. But though water is by far the most important and common of liquids, there are many others in frequent use ; for example, ether and methylated spirits. Will the same substances dissolve in them as in water? Are those substances which are insoluble in water also insoluble in these liquids? Try.

**EXPT. 13. To try the Solvent Action of Ether on Tallow.**—We know from daily experience that fats and oils do not dissolve in water. But will tallow, for instance, dissolve in ether<sup>1</sup>?

Place a small lump of tallow in a test-tube. Add about 5 c.c. of ether. Shake and stir with a glass rod. The tallow is dissolved.

The experiment shows that the solvent action of ether is very different from that of water. It is probable, therefore, that other liquids will also be found which will dissolve substances though they may be insoluble in water (*see Exercises*).

**PROBLEMS AND EXERCISES****Solution**

1. Find what weight of each of the following substances can be contained in 100 c.c. of a solution in water, saturated at the ordinary temperature :—

Blue vitriol.

Bichromate of potash.

Chlorate of potash.

Chloride of potassium.

Alum.

2. Prepare solutions in water containing 50 grams per 1,000 c.c. of each of the following :—

Common salt

Bichromate of potash.

Nitre.

Alum.

3. Measure out 50 c.c. of each of the solutions made in Exercise 2, and find the weights of the dissolved substances.

4. Prepare a solution of common salt containing 58.5 grams per litre.

<sup>1</sup> *Caution*: Take care that no flame is near, since ether vapour is very inflammable.

5. Prepare a solution of silver nitrate containing 17 grams per litre.
6. From the solution made in Exercise 4 prepare a weaker solution containing 5.85 grams per litre.
7. From the solution made in Exercise 6 make another solution containing 5 grams per litre.
8. Compare the hardness of a sample of hard water before and after it has been boiled.
9. Find whether the following will dissolve in ether at the ordinary temperature :—

Paraffin oil.

Salt

Soap.

Shellac.

Iodine.

Water.

Methylated spirits.

Nitre.

10. Find the solvent action of methylated spirits on the substances given in Exercise 9.



## CHAPTER XV

### CRYSTALLISATION

**CRYSTALLISATION.**—In by far the larger number of cases the solubility of a substance in water is increased by increase of temperature. Consequently, in general, if hot water be saturated with a substance, some of this substance will separate out in the solid state when the solution is cooled. In many cases it will separate out in crystals.

**EXPT. 1. To recrystallise Nitre.**—Measure about 100 c.c. of water into a small flask and heat it. Weigh out about 60 grams of nitre. Powder the nitre and dissolve it in the hot water. Filter the hot solution into a beaker or basin, and allow it to cool gradually. When cold, pour off the solution from the crystals which have formed. Transfer the crystals to some filter-paper, and dry them by gently pressing them with other pieces of clean paper. What weight of crystals have you obtained?

**Preparation of large well-formed Crystals.**—Perfect crystals are hard to obtain. The conditions of their growth are rest, freedom from dust, slow cooling, and a solution which is not too strong.

**EXPT. 2.** Prepare a warm saturated solution of nitre in a test-tube. Divide it into two equal portions. Put one aside to cool slowly. Cool the other rapidly under the tap. Shake it and rub the inside of the test-tube with a glass rod. A fine crystalline precipitate falls. In the first portion larger pointed, needle-like crystals will slowly form.

**EXPT. 3.** Dissolve some crystals of Glauber's salt (sodium sulphate) in a very little water. Filter into a clean flask. Close the mouth of the flask with a plug of cotton wool to exclude

dust. Allow the solution to become quite cold, and avoid shaking it. It does not crystallise. Remove the plug. Let fall some tiny specks of solid Glauber's salt into the solution. Feathery crystals start from where the dust fell, and in a few seconds the flask contains a solid mass of crystals.

The experiment with the nitre has shown the advantage of slow cooling and the disadvantage of shaking. That with Glauber's salt shows that dust should be excluded, and that from a strong solution crystals form all massed together.

**EXPT. 4. To prepare good large Crystals of Alum.**—Prepare a strong hot solution of alum by dissolving 30 grams of alum in 150 c.c. of water. Cool rapidly to make sure that it is strong enough to deposit crystals. If a great many crystals are formed pour off the liquid with some only of the crystals. Heat again until the crystals are dissolved. Filter the liquid into an evaporating basin. Place a piece of paper over the basin, and put it away in a cool quiet place. Examine in an hour's time. If many crystals have formed the solution is too strong; if there are no crystals it is too weak. Suppose there are a few crystals. Pick out the three best. Transfer them with a glass rod to another basin. Pour the cold "*mother-liquor*" from which they were deposited gently over them; and put aside in a warm place. Arrange them so that they are not in contact. Try also suspending a crystal by a hair looped around it in some of the mother-liquor. After twenty-four hours examine them again. Remove any tiny crystals adhering to them. Turn them over so that they may rest on different faces. If the alum solution is able to evaporate slowly it will go on enlarging these crystals. Notice that they are of an entirely different form from that of the crystals of nitre.

The shape of a perfect alum crystal is that of the regular solid called the *octahedron*, bounded by eight faces, each an equilateral triangle (*see* Fig. 66). In less perfect specimens two of these faces are often enlarged at the expense of the others. A model of the shape shown in Fig. 68 may be made by first cutting out a regular octahedron in soap and then cutting parallel slices off two opposite sides, as indicated in Fig. 67.

Almost every chemical substance has a certain shape or shapes in which it can crystallise—*e.g.*, crystals of salt are cubes, those of calcspar rhombs, and those of nitre six-sided prisms.

**WHAT IS A CRYSTAL?**—But what is the distinguishing feature of a crystal? How does a crystal differ, for example, from a piece of cleverly cut glass? The following experiment may suggest an answer.

**EXPT. 5. To show a difference between a Crystal of Rock-Salt and a Piece of Glass.**—Take a good crystal of rock-salt. Break it roughly with a hammer or in a strong mortar. Examine the fragments. They are all either little cubes or portions of cubes. Now break down a piece of glass rod in the same way. The fragments are of all shapes, and they appear to have no relation to one another.

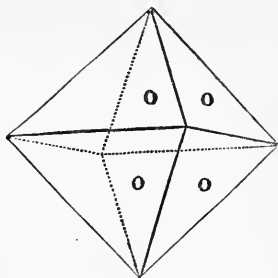


FIG. 66.—PERFECT CRYSTAL.

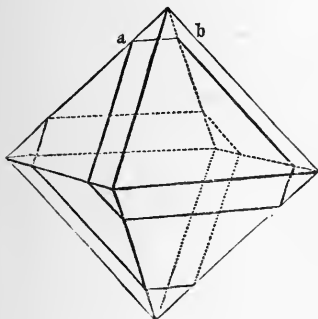


FIG. 67.

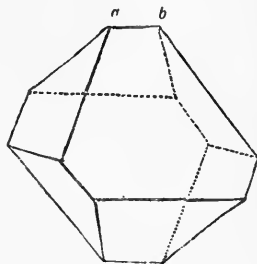


FIG. 68.

#### EXERCISE.

Crush a crystal of calspar, and examine the fragments.

The crystal of rock-salt seems then to possess definite lines along which fracture takes place, while the glass is broken as readily in one direction as in another. These lines along which

the crystal can be readily broken or cut are called *lines of cleavage*. A characteristic of a crystal seems to be that it possesses a *definite structure*.

The existence of this definite structure is also impressed upon us by the fact that all the crystals of a given substance belonging to the same general form, for example, crystals of calcspar, which are rhombs, have the *corresponding angles equal*.

**EXPT. 6. To show that the Corresponding Angles of Crystals of Calcspars are Equal.**—Cut two slips of card-

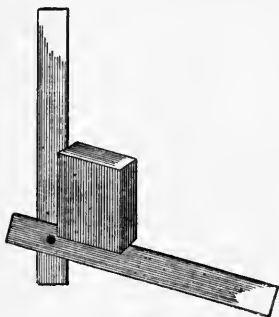


FIG. 69.

board. Fasten two ends together by piercing a pin through them, so that they can turn about the pin as a pivot. Choose two good crystals of calcspars. Place them on the bench in the same relative position. Take up one of the crystals, and place it in the angle between the two slips of cardboard, so that each slip lies along a face of the crystal (Fig. 69). Hold the slips in this position. Now take up the other crystal, and see whether the *corresponding*

angle in it will fit exactly into the angle between the slips of cardboard. It is found to be so.

### EXERCISES

1. Compare other corresponding angles on the same two crystals of calcspars.
2. Compare a third crystal of calcspars with one of those already taken.
3. Compare corresponding angles of crystals of
  - (a) rock-salt.
  - (b) alum.
  - (c) chrome alum.

It is found that however the general shape of a crystal may be distorted *the corresponding angles remain constant*.

This is well shown in snow-crystals, which assume many fantastic shapes, but they are all built up from a simple star whose rays intersect at  $60^{\circ}$ .

The thought will readily occur that it is the *structure*, spoken of above, which determines the outward form. And it might be supposed that a cube of rock-salt consisted of a very large number of very small cubes; but this is not necessarily the case.

**EXPT. 7. To build Geometrical Shapes with a Number of Spherical Balls.**—(1) Take a number of five-balls or marbles, and pack them in a square so that one ball is touched by four others. To prevent the balls rolling, enclose a square by a cardboard rim fastened at the corners. Pile others in the hollows between the balls. Then add yet another layer of balls, and so on. Finally, half a regular *octahedron* is the result.

(2) Arrange the balls in a triangle, the inner ball being surrounded by six others. Place balls in the hollows as before. The result is the corner of a *cube*.

A cubical crystal is, therefore, not necessarily built up with cubes.

On the other hand, it may be stated that there is every reason to believe that the forms of crystals *are* due to the symmetrical arrangement of the tiny particles which compose them.

## SEPARATION AND PURIFICATION OF SUBSTANCES BY MEANS OF CRYSTALLISATION.

—We have found that substances differ greatly in solubility in water. It follows, therefore, that many mixtures can be entirely separated into their constituents by treatment with water. This will be most complete if one portion of a mixture is soluble in water while the rest is quite insoluble.

**EXPT. 8. To separate a Mixture of Sand and Nitre.**—*i. e.*, to obtain from it all the sand dry and free from nitre, and the nitre dry and free from sand.

Grind some sand and nitre together in a mortar. Place about 10 grams of the mixture in a beaker. Add some hot water. Boil. The nitre dissolves, while the sand remains undissolved. The finer particles of sand are floating in the liquid, but would in time settle. We could then pour off or *decant* into a

second beaker a liquid which would be a solution of nitre, muddy with the finest sand, leaving in the first beaker sand wet with a solution of nitre. This would be a rough but not a complete separation. The sand must be washed, and the solution of nitre filtered.

Filter the muddy liquid. It is probably still hot, which is an advantage, since a liquid filters most quickly while hot.

(1) *To recover the Nitre from the Solution.*—If we only require *some* of the nitre, we need not wait for it all to run through the filter. Pour the clear nitre solution into an evaporating basin. Evaporate on a sand-bath, or still better on a water-bath, until the nitre is quite dry.

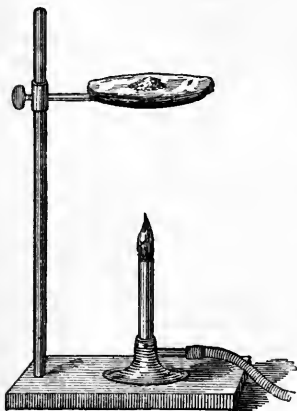


FIG. 70.

(2) *To clean the sand,* which was left wet with the nitre solution. Pour a little hot water on the sand in the beaker. Shake. If we wish to get either all the nitre or all the sand, then this washing must be run through the filter-paper. Wash similarly several times, transferring the sand to the paper. Some of the sand clings obstinately to the side of the beaker. Use the jet of the wash-bottle to dislodge the sand,

and wash it completely on to the filter. Wait till the liquid has completely run through the paper. Evaporate the filtrate, and add the residue to the nitre obtained before. Gently remove the filter-paper from the funnel. Open it out flat and place it on a pile of four filter-papers. Place the pile on wire gauze on a retort-stand ring about six inches above a small flame to dry slowly (Fig. 70).

But even when both constituents of a mixture are soluble in water it is possible to separate them, provided they are soluble in different degrees.

**EXPT. 9. Separation of Nitre from Common Salt—two Substances which are both Soluble in Water.**—Weigh

out 40 grams of nitre and 40 grams of common salt. Grind them together in a mortar with 100 c.c. of cold water. They do not entirely dissolve. If we refer to Fig. 65, p. 114, it is seen that the salt dissolves the more readily of the two in cold water, so that what remains undissolved is chiefly nitre. Pour the whole upon a filter-paper and thus separate the undissolved substance from the solution. Keep the filtrate: call it *a*. Pour a very little more cold water onto the filter-paper, so as to wash away the solution. Open out the filter-paper, and transfer the substance on it to a small beaker. Add about 20 c.c. of cold water. Heat it gradually until the substance is dissolved. Cool the beaker by placing it in a larger one containing cold water. After a time crystals of nitre separate out. The small amount of common salt that was left with the nitre upon the filter-paper remains now almost wholly in solution. Pour off the solution from the crystals of nitre. Dry them on filter-paper. Recrystallise again from a small amount of water.

We have thus first separated a crude nitre, and secondly by *recrystallising* this we have purified it from small amounts of common salt. This is an example of the way in which substances may often be purified from foreign material.

#### PROBLEM

Boil down the filtrate *a* obtained in the last experiment for a few minutes. Filter off what separates out, and examine it.

**Water of Crystallisation.**—Try the following interesting experiment.

**EXPT. 10. The effect of Heat on Epsom Salts.**<sup>1</sup>—Weigh a small porcelain crucible. Add 1 to 2 grams of Epsom salts. Weigh. Heat the crucible gently on a sand-bath. The crystals melt and the liquid then begins to boil. Vapour is evolved, which has no odour. Hold an inverted dry funnel in the vapour. A dew is formed on the cool glass. It appears to be water. Go on heating until no more vapour is given off. Cool. Weigh. The solid has lost weight. Repeat, using a lid with the

<sup>1</sup> Magnesium sulphate,

crucible, and heat the crucible with the lid on (slightly tilted, Fig. 71) for fifteen minutes. In an experiment,

0.875 grams Epsom salt lost 0.431 grams,

hence, 100           ,,           ,,            $\frac{.431}{.875} \times 100$  or 49.2 grams.

Do you find the same loss ?

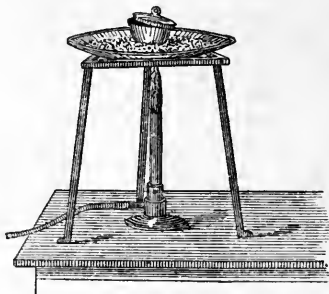


FIG. 71.

**EXPT. II. The Action of Water on the Residue left after Heating Epsom Salts.**—Try the solvent action of water on the residue obtained in the last experiment. It is readily dissolved. Will it be recovered if the solution is evaporated? To test this we must work with more material.

Heat about 10 grams of Epsom salts, just as in Experiment 2. Dissolve the residue in as little water as possible. Evaporate the solution on a water bath. What is obtained? Crystals of Epsom salts are obtained again, and not the non-crystalline residue.

Then it seems that the crystalline form has been recovered, because water has been added. Moreover, the vapour which was given off when the crystals were heated must have been water vapour. Water which can be driven off by heat in this way from crystals is called *water of crystallisation*, and generally the crystals can be reformed by the addition of water. We shall find that there are many solid crystalline substances which contain such water.



## PROBLEMS AND EXERCISES

## Crystallisation

1. Prepare crystals of the following substances :—Rochelle salt, chrome-alum, copper sulphate, ferrous ammonium sulphate, salt, nitre.
2. Find what happens when crystals of chrome-alum are placed in a crystallising solution of common alum.
3. Make mixtures of the following substances, and then separate samples of the constituents by means of water :—

Chalk and common salt.

Nitre and charcoal.

Sand and sugar.

4. Mix thoroughly 40 grams of chlorate of potash with (a) 40 grams potassium chloride, (b) 40 grams bichromate of potash. Then separate by crystallisation a sample of pure chlorate of potash from each mixture.
5. Find the percentage weight of water of crystallisation in 100 grams of the following substances : A, B, C, D.<sup>1</sup>

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<sup>1</sup> *Note for the Teacher.* The following are suitable substances :  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ ,  $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$ ,  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ,  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ .

## CHAPTER XVI

### RAW MATERIALS

**The Earth is the Chemist's Storehouse.**—The most careless observer cannot fail to conclude that the soil and rocks which form the crust of the earth are not everywhere of the same nature. The purple hue of Welsh slate, the transparency of Derbyshire spar, the sombre gray of Scotch granite, or the yellow sands of the sea-shore point to differences of material. The early alchemists understood this, and stored the different minerals and ores they found in the crust of the earth upon the shelves of their laboratories. They made experiments with them in their crucibles and alembics, and succeeded in producing many new substances.

So, too, for us the earth is a storehouse, and we have but to search it for raw material. We may obtain common salt from Northwich, chalk from Dover, iron ore from Spain, or saltpetre from India. With these and other native or raw materials we may then make all manner of experiments and hope to prepare new substances.

It will be wise, therefore, in the first place to obtain some of these materials, and proceed to examine them, noting points of resemblance and difference. We must carefully observe their appearance, forms and their colours, and it will perhaps be useful to try the effect of heat upon each; then when a mineral is brought to us to be named we shall often be able to identify it.

## COMMON SALT.

Common salt is always present in sea water along with other substances, and for ages men have obtained it from the waters of the oceans. If sea water be evaporated, cubical crystals of common salt separate out. These must be purified by re-crystallisation. In warm climates the evaporation is often done in great basins on the sea-shore, and the crystals of salt are raked off. "Bay salt," which may be purchased through a chemist, shows the cubical crystals.

Common salt also occurs as huge deposits of rock-salt in various strata of the earth. The chief of these in England is at Northwich in Cheshire, where there are two great beds of rock-salt, together nearly 60 feet in thickness. It is probable that this rock-salt was deposited ages ago by the evaporation of sea water. A common method of bringing the salt to the surface from these buried stores, is to bore down to them and to pour water into the boring. After lying in contact with the salt beds the water becomes thoroughly saturated. The rich brine is then pumped up to the surface and evaporated. One hundred parts by weight of brine will yield about 23·5 parts of salt.

**EXPT. 1. Common Salt is not changed by Heat.**—Place about 2 grams of common salt in a porcelain crucible. Place the crucible in a sand-bath upon a tripod stand. Heat it with a Bunsen burner. The salt is not changed in appearance. If the salt be pure no vapour is given off. Hold an inverted funnel over the crucible. Frequently common salt is moist ; if so, a dew of water will be deposited on the glass.

## CHALK.

Chalk is found right across England from Flamborough Head, along the Wolds of Yorkshire and Lincolnshire, past Newmarket and the Gog Magog Hills in Cambridgeshire, to Salisbury plain. Thence it stretches in the North and South Downs to Dover and Beachy Head. Chalk has been shown to consist largely of the broken fragments of tiny shells or foraminifera, such as are now forming the gray ooze or soft

bed at the bottom of the Atlantic, on which the telegraph cables rest.

**EXPT. 2. To find the Tiny Shells in Chalk.**—With a camel's hair brush dust the surface of a lump of chalk into a watch glass containing some water. Carefully pour off the water from the watch glass, and then fill it up again with water. In this way wash the dust again and again with water. The heavier particles sink to the bottom, the lighter are swept away. Examine the residue with the microscope. Real chalk must be used for this experiment, not the soft "chalk" used on blackboards.

For further information the student is recommended to read Huxley's "Physiography" on this subject.

**Limestone** is the general name given to many stones which when burnt in a kiln yield lime. To include all such stones we ought to call *chalk* and *marble* limestones. A large part of the Pennine Range consists of hard dark blue-gray stone streaked with white called "mountain limestone." It contains numerous fossils of shells and corals, suggesting that it has once been mud at the bottom of a coral sea. It is found from Cheviot past Crossfell and Ingleborough to the Peak. The streams which flow over it, like those in chalk districts, are hard. Their waters dissolve the rock into fantastic caverns.

Another sort of limestone called *Dolomite* or *Magnesian Limestone* extends in a narrow band from the coast of Durham past Knaresborough and Pontefract in Yorkshire, to Nottingham.

**EXPT. 3. The Effect of Heat on Chalk: Formation of Quicklime.**—Heat a lump of chalk for some hours in the centre of a hot coal fire. Remove it, and allow it to cool. Taste a very small portion. It is very caustic.

It has long been known that when chalk or limestone is exposed to a violent heat, it is converted into this caustic substance, called *quicklime*.

In chalk and limestone districts you may have seen the kilns in which quicklime is obtained by heating a mixture of chalk or limestone with a little coal.

#### SAND, FLINT, AND QUARTZ

Sand, flint, and quartz all resemble each other in being extremely hard—hard enough to scratch glass. Various rocks

called *sandstones* consist of grains of sand cemented together. Such are the sandstones of the coal measures, the millstone grit of the Yorkshire moors, the red sandstone of Devonshire. Pure sand is used for making glass. Flint nodules are found in chalk. Flint was used for arrowheads by prehistoric man. Quartz is one of the constituents of granite.

EXPT. 4. **To powder Flint.**—Heat some flints to redness and drop them into cold water. Examine a piece. It has become opaque and brittle, and can be readily powdered.

### WASHING SODA.

Up to the beginning of the present century the seaweeds cast up on the coasts of Normandy, Scotland and Ireland were the only source of soda. The seaweeds were burnt, and their ashes—called *kelp*—were dissolved in hot water. On evaporating the liquor crystals of soda were obtained. It is now made from common salt in very large quantities and by more complicated processes.

Shake a few soda crystals with water in a test-tube. They dissolve extremely easily. Place a drop or two of the solution on the palm of the hand, and rub with a finger. It produces a soap-like sensation. Hence, the substance is called *washing soda* because it is found to have the same cleansing action as soap, but it is sometimes simply called *soda*.

Taste a crystal. It has a peculiar *alkaline* taste.

Find the effect of heat on crystals of washing soda. Proceed exactly as in Experiment 10, p. 123. Do you not find that the crystals seem to contain a good deal of water of crystallisation? How much?

### POTASH.<sup>1</sup>

When land plants are burnt an ash is left which contains a substance very similar to soda. This used to be done in large pots, and so the ash was called *pot-ash*. The ash was then treated with hot water, and the solution so obtained was filtered from any undissolved matter. On evaporating the solution, a brown mass of crude potash was obtained. To

<sup>1</sup> Potassium carbonate.

purify it, it was heated strongly, and then re-treated with water.

Potash is a white granular non-crystalline powder. Taste a little. It has the same alkaline taste as soda.

Find the effect of heat on potash in the same way as in Experiment 10, p. 123. Does it melt? Is much vapour given off?

### SALTPETRE OR NITRE.

In certain parts of India, the soil, after the rainy season, becomes covered with small crystals of a different salt from any referred to so far in this chapter. This "salt of the rock" was called *sal petrae* by the alchemist Geber, and is known at the present day as *saltpetre* or *nitre*. To separate this substance, the soil is scraped and treated with water, in which the nitre readily dissolves, and on the evaporation of this solution the nitre crystallises out.

Taste a small crystal of nitre. It has a bitter cooling taste.

Examine and describe carefully the appearance of nitre. Make a drawing of a good crystal.

**EXPT. 5. The Effect of Heat on Nitre.**—Place about 1 gram of nitre crystals in a test-tube. Heat in a Bunsen flame. The crystals melt, and the liquid seems to boil, and a number of little bubbles are evolved. Allow to cool. Try to recrystallise from water. You probably will find that you cannot recover any crystals of nitre. It seems then the heat has changed the nitre into a different substance.

### IRON PYRITES.

Iron pyrites occurs very widely in many different places. When pure it has a brass-like colour. Small quantities of it may frequently be found in lumps of coal.

Examine a piece carefully. Perfect crystals of iron pyrites take the forms of the regular solids; the cube and the octahedron are common forms.

**EXPT. 6. The Effect of Heat on Iron Pyrites.**—Heat a small quantity in a test-tube. A sulphurous vapour is given off, and a yellow deposit of sulphur is formed in the cooler part of the tube, while the mineral acquires a deep red colour.

GREEN VITRIOL.<sup>1</sup>

When iron pyrites is long exposed to the air and moisture, a greenish liquor is obtained. If the liquor is evaporated, large green glass-like crystals form. These were known to the alchemists under the name of *green vitriol*.<sup>2</sup>

At the present time large quantities of green vitriol are obtained in this way from the pyrites occurring in the coal of South Lancashire.

**EXPT. 7. The Effect of Heat on Green Vitriol.**—Place a few crystals in a test-tube. Heat cautiously in a flame. It melts, boils, and a dew of water is deposited on the cool upper part of the tube. Soon it becomes dry and a hard white mass is left. Then heat strongly. A sulphurous vapour is given off, and the mass becomes red, where it is in contact with the glass.

## BLUE VITRIOL.

The drainage water from copper mines has frequently a blue colour. If such water is partially evaporated and the liquid is then allowed to cool, fine blue crystals are obtained. These were known to the alchemists as *blue vitriol*.

**EXPT. 8. The Effect of Heat on Blue Vitriol.**—Repeat Experiment 10 p. 123, but do not use a very large flame. Observe that the substance becomes gradually colourless. Find the percentage loss of weight.

Do your results agree with the following?—

1·192 gram blue vitriol lost	0·433 gram	=	36·3 per cent.
2·102        „        „	0·770        „	=	36·6        „
3·136        „        „	1·146        „	=	36·5        „
		mean =	36·5        „

SAL-AMMONIAC.<sup>3</sup>

A remarkable salt was prepared long ago in Egypt from camel's dung. When the dung was heated a heavy soot was

<sup>1</sup> Sulphate of iron.

<sup>2</sup> Latin *vitrum*, glass.

<sup>3</sup> Ammonium chloride.

deposited, from which the salt called *sal-ammoniac* was made. In those days, just as at the present time, it was widely used as a medicine.

Later, it was prepared from the soot obtained when such animal refuse as hoofs, horns and hair were heated.

**EXPT. 9. The Effect of Heat on Sal-Ammoniac.**—Cover the bottom of a porcelain basin with a layer of sal-ammoniac. Invert a glass funnel over it. Place the whole on a sand-bath over a burner. Heat gently. White fumes are given off, and a white deposit is formed on the sides of the funnel. Finally, the whole of the sal-ammoniac disappears from the basin and is found again unchanged on the funnel, and if some of the deposit is heated it behaves again in exactly the same way as the original sample. This is an example of what is called *sublimation*.

## PROBLEMS AND EXERCISES

### Raw Materials

1. Describe fully a piece of slate.
2. Describe fully a lump of granite.
3. Describe fully the appearance of the substances, A, B, C, D, E, F.<sup>1</sup>
4. Find the effect of heat on the substances, G, H, I, J, K, L.<sup>2</sup>
5. Identify the substances, M, N, O, P, Q, R.<sup>3</sup>
6. Find whether any water is contained in the substances, V, W, X, Y, Z.<sup>4</sup>
7. Find what occurs when potash is liquefied by exposure to the air.
8. Find whether crystals of blue vitriol can be recovered from the white residue obtained on strongly heating the crystals.
9. Find the solubility in water of (i) chalk, (ii) quicklime.
10. Mix thoroughly 40 grams of common salt and 40 grams of Epsom salts. Endeavour to separate a sample of each by means of crystallisation from water.

<sup>1</sup> *Note for the Teacher.* The following are suitable substances, *e.g.* HgO, HgI, Pb<sub>3</sub>O<sub>4</sub>, NaCl, CaCO<sub>3</sub>, NH<sub>4</sub>NO<sub>3</sub>.

<sup>2</sup> *e.g.* NH<sub>4</sub>Cl, NaHCO<sub>3</sub>, Pb(NO<sub>3</sub>)<sub>2</sub>, KNO<sub>3</sub>, ZnSO<sub>4</sub>.7H<sub>2</sub>O, KClO<sub>3</sub>.

<sup>3</sup> *e.g.* NaCl, KNO<sub>3</sub>, CuSO<sub>4</sub>, CaO, NH<sub>4</sub>Cl, Na<sub>2</sub>CO<sub>3</sub>.

<sup>4</sup> *e.g.* Na<sub>2</sub>SO<sub>4</sub>.10H<sub>2</sub>O, Rock Salt, KNO<sub>3</sub>, K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>, ZnSO<sub>4</sub>.7H<sub>2</sub>O.



## CHAPTER XVII

### THE PREPARATION OF THE COMMON ACIDS AND ALKALIES

THE Alchemists made numberless experiments with the raw or native materials which they found ready for them in the crust of the earth, and learnt how to prepare many new substances. Amongst these none have proved of greater importance and have found more frequent use than certain acid and alkaline liquids.

We also must now learn how to prepare these liquids for future use, and how to distinguish one from another, by such simple tests as we can discover.

#### A. THE COMMON ACIDS.

##### 1. Oil of Vitriol.<sup>1</sup>

**EXPT. 1. Preparation of Vitriolic Acid from Green Vitriol.**<sup>2</sup>—Weigh out about 10 grams of green vitriol crystals into an iron dish or porcelain basin. Heat over a flame, stirring with a glass rod until all water is expelled and a brittle white mass is left. Break the mass into small lumps (but do not powder it). Introduce these into a small retort *a* (Fig. 72). Support the retort so that it can be heated by a naked flame. Let the neck dip into a little water in a test-tube *b*. Heat the retort at first

<sup>1</sup> Sulphuric acid.

<sup>2</sup> Lecture Table Experiment.

with a small flame, and finally with the full flame for at least fifteen minutes. Remove the test-tube. Then turn the flame down gradually, so that the retort may cool slowly.

Is the water in the test-tube still merely water? Pour two or three drops into another test-tube. Taste these. The liquid has a very acid taste. This acid liquid, thus prepared from green vitriol, was called by the early chemists *vitriolic acid*. It is perhaps the most useful of all the acids that are known.

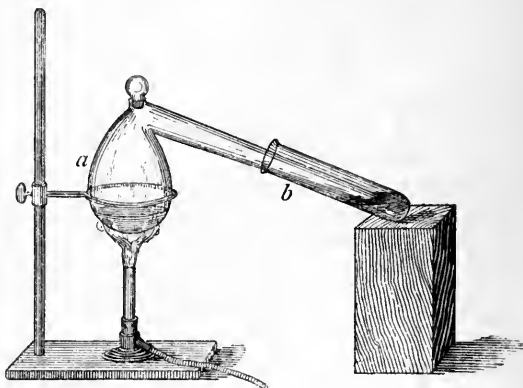


FIG. 72.

A better method of preparation has been discovered since the time of the alchemists, but this is not the place to describe it.

Transfer the residue in the retort *a* on to a sheet of white paper. Observe that those portions which were most strongly heated have acquired a deep red colour. This substance when mixed with oil is used largely as a red paint. Jewellers' rouge also consists of this material.

**EXPT. 2. Vitriolic Acid when heated gives off dense White Fumes.**—Pour some of the acid prepared in experiment 1 into a small porcelain crucible. Heat it over a small flame. It boils down to a small bulk giving off steam. Then the fumes become thick and white and possess a pungent odour, and the remaining liquid has an oily appearance. On

this account the acid is frequently called *oil of vitriol*. Another name for it, much used now, is *sulphuric acid*.

**EXPT. 3. The Mixing of Oil of Vitriol and Water.**—Three-quarters fill a beaker with water. Mark the level of the water on the side of the glass. Measure into another beaker about one-eighth of this volume of acid. Pour some of the acid into the water in a thin stream. It sinks to the bottom; so the acid is heavier than water. Stir the water round with a glass rod. Add the rest of the acid in the same way. Grasp the beaker in the hand. The mixture has become quite hot, and steams.

Oil of vitriol when mixed with water becomes sometimes hotter than the boiling point of water. What would happen if a little water were poured on to strong oil of vitriol? Being lighter, the water would float on the surface. The liquids would mix at the bounding surface and become very hot, hot enough to turn some of the water to steam, and throw the acid about with explosive violence. Hence, *never pour water on to strong oil of vitriol*.

**EXPT. 4. To show the Destructive Action of Oil of Vitriol.**—(a) Write a word with some dilute oil of vitriol with a glass rod on a piece of *paper*. Dry the paper thoroughly by gently wafting it above a Bunsen flame. The paper is charred where the acid was traced.

(b) Place about 10 grams of lump *sugar* in a tall beaker or jar. Add about 10 c.c. of hot water. Place the beaker on a plate or dish. Then add at once about 10 c.c. of strong oil of vitriol. The mixture blackens at once, and froths up.

(c) Mix 1 or 2 c.c. of oil of vitriol with about 10 c.c. of water. Pour the mixture on a small piece of zinc in a beaker or basin. The zinc is dissolved.

*Note on Sulphuric Acid.*—(1) Should any acid be spilt, throw on some quicklime or powdered chalk, and mop up with an old duster. (2) Never let the acid boil; the fumes are very disagreeable.

After the discovery of oil of vitriol the alchemists tried the effect of heating it with many substances, with the result that they were led to discover other acid liquids, as will be described in the course of the following pages.

## 2. Spirit of Salt.<sup>1</sup>

Another acid liquid was obtained by the alchemists by heating common salt with certain substances. Glauber, who lived at the beginning of the seventeenth century, first prepared it from common salt and oil of vitriol, and it is always made at the present time from these.

**EXPT. 5. Preparation of Spirit of Salt.**—Weigh out about 30 grams common salt. Introduce into a litre flask.

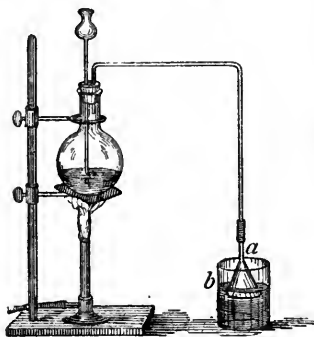


FIG. 73.

Attach a cork fitted with a thistle-funnel, and a delivery tube bent twice at right angles. Attach an inverted funnel *a* to the end of the delivery tube by a small piece of rubber tubing (Fig. 73). Support the flask upon wire gauze on the ring of a retort-stand, and place the funnel within a small beaker *b*. Add enough water to the beaker to cover the mouth of the funnel.

Measure out 30 c.c. strong oil of vitriol. Pour it little by little down the thistle-funnel

on to the salt. Fumes are given off, and bubbles are forced through the water. As soon as all the acid has been added, heat the flask gently. Soon the bubbles cease to come through the water.<sup>2</sup> After a time the water in the beaker begins to fume strongly. As soon as this is the case, remove the beaker and also the flame.

Has anything passed into the water in the beaker? It seems probable from the way in which it fumes. Dip a glass rod into the water and place a very small drop upon the tongue. It is very acid, and has a different taste from oil of vitriol. Where has the acid come from? The alchemists thought that it was a

<sup>1</sup> Hydrochloric acid.

<sup>2</sup> Care must be taken that the water in the beaker does not rush back along the tube into the flask. If it begins to do so, at once heat more strongly or remove the beaker.

spirit which the oil of vitriol had driven out of the salt. They therefore called it *spirit of salt*; but now the fuming acid liquid is called *hydrochloric acid*. Can the acid be driven off again from the water?

EXPT. 6. **Spirit of Salt is Volatile.**—Pour about 5 c.c. of the acid into a porcelain dish. Heat it gently over a flame. Acid vapours are given off it and it is soon completely evaporated.

### 3. Spirit of Nitre.

The alchemist Geber, who lived about the twelfth century, described an acid liquid obtained by heating nitre with green vitriol. Since the acid liquid was obtained by means of nitre it used to be called *spirit of nitre*. It is now usually termed *nitric acid*.

Although spirit of nitre can readily be obtained by Geber's method, it is now usually prepared by heating nitre with oil of vitriol in place of green vitriol. Glauber first prepared it in this way, and he called it "spiritus nitri fumans Glauberi" or "Glauber's fuming spirit of nitre."

EXPT. 7. **Preparation of Nitric Acid from Nitre and Oil of Vitriol.**—Weigh out 20 grams of nitre, and measure out 30 c.c. strong oil of vitriol. Support a small retort *a* on wire gauze or as shown in Fig. 74. Pass the neck into a small dry flask *b*. Support the flask *b* under a tap or in a trough, and keep it cool with running water.

Slide the nitre into the retort, and pour the oil of vitriol in through a funnel, taking care that neither gets down the long tube of the retort. Replace the stopper of the retort. Heat gently. The nitre melts, and the liquid begins to boil. A ruddy vapour rises, and drops form and run down the tube. A yellowish liquid collects in the flask. Keep the liquid just boiling, but do not let the temperature rise unnecessarily. As soon as the liquid in the retort begins to thicken, remove the flame.

Pour the remaining liquid from the retort into any empty evaporating basin. Put it aside. On cooling it hardens into a solid mass. If this had been left in the retort it would have been rather difficult to get it out.

Can nitric acid be evaporated just as spirit of salt (hydrochloric acid)?

**EXPT. 8. Nitric Acid is Volatile—Tests for Nitric Acid.**—(1) Pour about 5 c.c. of the liquid into a porcelain dish. Heat gently until the liquid is evaporated. Vapours are evolved possessing a peculiar odour quite unlike the fumes of oil of vitriol or of spirit of salt.

(2) Place a small drop of the liquid upon a finger nail, and another upon the skin of the hand. Wash them off after a few seconds. Bright yellow stains are left.

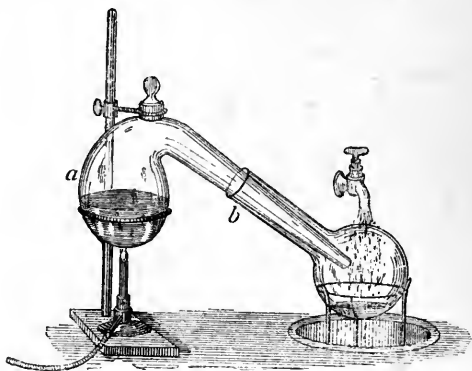


FIG. 74.

(3) Place a small piece of copper in the liquid. Red fumes are at once evolved, and the liquid becomes blue.

Show that oil of vitriol and spirit of salt do not attack copper in this way. On account of the corrosive action of nitric acid it is used frequently to be called *aqua fortis* (i.e., strong water).

#### 4. Vinegar or Acetic Acid.<sup>1</sup>

Acetic acid is another acid which is frequently used, and therefore may be mentioned here, although it is not prepared from any materials which are found native in the rocks or crust of the earth.

<sup>1</sup> Omit in a first study of the subject.

When beer or a weak wine such as claret is left exposed to the air, it quickly becomes sour. This is because the alcohol has been changed into the acid called acetic acid or vinegar.

The best way to convert weak wine into vinegar is the German "quick-vinegar process." A large cask with perforated sides is filled with beech wood shavings, which are moistened with a little vinegar. Weak wine is then poured in at the top, and it trickles slowly through the shavings, and the liquid which collects at the bottom is again and again poured down until all the alcohol is changed into vinegar. A kind of fungus called *mycoderma aceti* is found on the shavings, and it is this which in the presence of air effects the change.

**EXPT. 9. Acetic Acid is Volatile.**—Heat about 5 c.c. of the acid in a basin. Observe that the vapour has a very pungent odour, which is quite different from that from any of the other acids. No residue is left.

## (B.) SOME COMMON ALKALIES.

We must now learn a little about some substances very different in character from the acids, which the alchemists also knew how to prepare.

### 1. Caustic Soda.

It was found long ago that if a solution of washing soda was heated with lime the solution changed in character and acquired a very caustic nature. You will have an opportunity of trying this yourself later (see page 278, Expt. 15).

If such a solution is allowed to stand till clear, and is then evaporated, a white non-crystalline solid is obtained. A piece of this solid or of the strong solution will quickly cause a blister when placed upon the skin. On this account it is called *caustic soda*, and must be handled with care.

Obtain a piece of solid caustic soda. Examine it and see how readily it dissolves in water.

Dilute with a large amount of water a little of a solution of caustic soda, such as is in a bottle on your shelf. Taste it, and notice its soapy "alkaline" nature. Wet your fingers with the solution: they feel soapy and slippery. Next time your hands

are greasy with oil, try washing them in a dilute caustic soda solution. They will become beautifully clean. A stronger soda solution would remove the skin as well as the dirt.

## 2. Caustic Potash.

Another caustic solution is formed if a solution of potash is heated with lime. If the clear caustic solution is evaporated, a white solid is obtained which is almost exactly similar in appearance to caustic soda. Moreover its solution in water has a very similar caustic and alkaline nature.

This substance is called *caustic potash*. It is usually sold in white sticks.

## 3. Volatile Spirit of Sal-Ammoniac or Ammonia.

EXPT. 10. **Preparation of Spirit of Sal-Ammoniac or Ammonia.**—Weigh out 10 grams sal-ammoniac. Powder it.

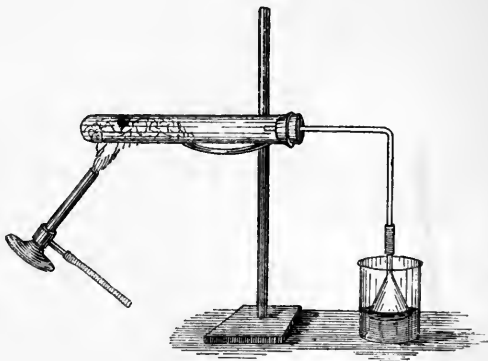


FIG. 75.

Dry it by heating it gently in a basin on a sand-bath. Allow it to cool. Meantime weigh out 15 grams quicklime, and powder it. Mix the two together. Put the mixture in a large test-tube. Fit the tube with a cork and a delivery tube bent once at right angles. Attach an inverted funnel to the delivery tube. Support the test-tube on the ring of a stand, so that the



funnel dips into a small beaker half filled with water, with its mouth just below the surface of the water, as shown in Fig. 75.

Heat cautiously by wafting the flame about under the tube. At first bubbles of air are forced up through the water, and then these cease. Heat 5 minutes longer. Take care that the water in the beaker is not sucked into the test-tube. Then remove the beaker, and afterwards the flame.

Observe the pungent and peculiar odour the water in the beaker has acquired. Since it was acquired by heating sal-ammoniac with lime, the early chemists called it the *volatile spirit of sal ammoniac*. It is now usually called *ammonia*.

Ammonia can also be prepared by heating shavings of horn. Indeed, this method of preparation was at one time so commonly used that the pungent liquid came to be known as "spirits of hartshorn," a name which is still common for it.

**EXPT. 11. Tests for Ammonia.**—(1) Pour a little of the solution into a porcelain crucible. Heat over a small flame. It is quickly evaporated, and the vapours possess the characteristic odour already observed.

(2) Dip a glass rod in strong hydrochloric acid. Hold it in the vapours arising from the solution. Dense white fumes are formed.

Repeat with a rod dipped in sulphuric acid and in nitric acid.

#### 4. Lime Water.

**EXPT. 12. Slaked Lime is somewhat Soluble in Water.**—Half fill a test-tube with water. Add some powdered slaked lime. Shake. The liquid becomes white and milky. The lime has not all dissolved. Perhaps some of it has dissolved. Allow the test-tube to stand. The white mud, which is only in suspension, begins to settle. In time it will all settle and leave the liquid quite clear. To save time filter the liquid. Evaporate a drop on platinum foil. A white stain is left showing that some of the quicklime was dissolved.

Quicklime then is slightly soluble in water. The solution is called *lime water*. Observe its alkaline taste. Lime water is generally prepared by shaking quicklime with water and allowing

it to settle. The shaking is repeated to strengthen the solution. The clear liquid is then decanted or syphoned off.

The acid and alkaline solutions we have been studying are commonly spoken of as *acids* and *alkalies*.

### PROBLEMS AND EXERCISES

#### The Common Acids and Alkalies

1. Find the effect of evaporating the liquids A, B, C, D, E,<sup>1</sup> in a porcelain basin.

2. You are given samples of the following pairs of liquids. Distinguish the members of each pair.

Oil of vitriol and nitric acid.

Hydrochloric acid and nitric acid.

Oil of vitriol and acetic acid.

Acetic acid and ammonia.

3. Find the effect of heating 1 part of nitre with 2 parts of green vitriol in a small retort the neck of which dips into a little water in a flask.

4. Find the effect of heating sal-ammoniac with oil of vitriol, as in Experiment 5.

5. Find the effect of heating Chile saltpetre with oil of vitriol, as in Experiment 7.

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<sup>1</sup> *Note for Teacher.*—e.g.  $\text{NH}_4\text{OH}$ ,  $\text{Ca}(\text{OH})_2$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{HCl}$ ,  $\text{HNO}_3$ ,  $\text{CH}_3\text{COOH}$ .

## CHAPTER XVIII

### THE ACTION OF ACIDS AND ALKALIES UPON LITMUS—THE FORMATION OF SALTS

Chemists found 200 years ago that the acid and alkaline substances we have just learnt to prepare had a curious effect on certain vegetable dyes. Of these dyes *litmus* is one of those most readily acted on, and it is therefore very suitable for our experiments. It is a solid blue substance which is extracted from certain kinds of lichen.

**EXPT. 1. To make an Infusion of Litmus.**—Weigh out about 5 grams of solid litmus. Grind it up in a mortar, and transfer into a flask. Half fill the flask with cold water and shake well. Allow the powder to settle, and then pour off the liquor. In this way certain impurities are washed out of the litmus.

Next add about 100 c.c. of water, and heat the flask nearly to boiling for a few minutes. Allow to settle (if possible for some hours in a warm place). Decant off the intensely blue solution from the sediment, and place the solution in a bottle.

**EXPT. 2. To find the effect of Acid and Alkaline Liquids on the Blue Litmus Solution.**—Let us write down the names of the acid and alkaline liquids we have on our shelves :—

*Acid liquids.*

Sulphuric acid.  
Hydrochloric acid.  
Nitric acid.  
Acetic acid.

*Alkaline liquids.*

Caustic soda.  
Caustic potash.  
Ammonia.  
Lime-water.

Try the effect of each of these liquids on the blue litmus solution in the following way :—Fill one-third of a test-tube with water ; add 5 drops of the litmus solution by means of a pipette or piece of glass tubing ; then add 5 drops of each liquid and stir. Carefully observe any change in colour. State the effect you notice in a table thus :—

Liquid taken.	Effect on the Blue Litmus solution.
Sulphuric acid.	Change to brick-red colour.

You have probably found that all the acid liquids turn the blue litmus solution *red*, while all the liquids possessing an alkaline instead of a sour taste leave the litmus solution *blue*, or render it still more blue. By this colour change, then, it seems that we can readily distinguish between an acid and an alkaline solution. But have no other liquids beside acid and alkaline solutions any colour effect upon litmus? This question must be tested by experiment.

#### PROBLEM

Find whether blue litmus solution is effected by alcohol, chloroform, milk, a solution in water of common salt, washing soda or of any other substance you can think of. State your observations in a table.

**EXPT. 3. To make Red and Blue Litmus Papers.**—Take a strong blue infusion of litmus, some very weak nitric acid, some very weak caustic soda solution, and a pipette. Divide the blue litmus solution into two parts. To one portion add a few drops of the dilute acid till it is just reddened. To the other portion add one drop of the weak caustic soda solution so that it has a blue colour. Soak pieces of blotting-paper in each solution. Hang them up to dry. Then cut them into strips, and keep in a glass-stoppered bottle.

## PROBLEM

Try the colour-effects of acid, alkaline and other liquids upon these red and blue litmus papers, by putting drops of the liquids upon strips of each kind by means of a glass rod. State what you observe in a table thus.

Liquid.	Effect on Red Litmus.	Effect on Blue Litmus.
Sulphuric acid.	No visible change.	Red colour.

**EXPT. 4. To try the Delicacy of the Litmus Test for an Acid.**—Add a cubic centimetre of strong sulphuric acid to a litre of water. Shake well. The acid is now diluted one thousand times. Dip in a blue litmus paper. It is reddened. Find whether the acid can still be detected when this weak solution is made 10 times and 100 times weaker still.

**EXPT. 5. To prepare a Purple or Neutral Litmus Solution.**—Take some blue litmus solution, some very weak acid solution, some very weak alkali, and a pipette. Drop the dilute acid into the blue litmus solution and stir with a glass rod until the colour is half-way between red and blue. Probably you will overstep the mark. If so, add the dilute alkali drop by drop very carefully. If the litmus changes in colour sharply from red to blue, weaken both the acid and alkaline solutions, and try again.

You will at last succeed in getting a *purple* or neutral litmus solution, intermediate in colour between red and blue.

To two portions of the purple or neutral litmus solution add acid and alkali respectively. One changes to red and the other to blue. Then, if we keep a stock of this one neutral solution, we can use it to detect both acids and alkalies, instead of using both red and blue solutions or papers.

Keep the neutral litmus solution in a glass-stoppered bottle.<sup>1</sup>

<sup>1</sup> If the solution loses its colour the colour will be quickly restored on removing the stopper and shaking the solution with a little fresh air.

## PROBLEMS

1. Prepare blue-purple and red-purple litmus solutions, intermediate in colour between blue and purple, and red and purple respectively.
2. Find whether lime-juice, soap, vinegar, washing soda, bicarbonate of soda, common salt, nitre, borax, milk, lead nitrate, &c., have any effect on neutral litmus solution.
3. Find the effects of acids and alkalis upon methyl orange, phenol phthalein, infusion of red cabbage, infusion of logwood, infusion of violet leaves.

## THE FORMATION OF SALTS

We have found that of the acids and alkalis known to us :—

1. An acid reddens blue litmus, and leaves red litmus unchanged.
2. An alkali turns red litmus blue, and leaves blue litmus unchanged.

What will be the effect of mixing an acid with an alkali in the presence of litmus ?

**EXPT. 6. To find the Result of adding Hydrochloric Acid to Caustic Soda until the Liquid is Neutral.**—Place about 10 c.c. of strong caustic soda solution in a basin. Add a few drops of litmus solution so that it is distinctly blue. Fill a pipette with dilute hydrochloric acid. Run the acid carefully drop by drop into the basin, stirring continually until the colour suddenly changes from blue to red. Then add some very dilute caustic soda solution until the liquid has a neutral or violet tint.

What is there now in the solution ? Is it simply a mixture of the acid and the caustic soda ? Suppose we find out what is left when the liquid is evaporated, for you will remember that hydrochloric acid, or spirit of salt as it used to be called, is volatile, so that we may expect to drive off the acid and have only the caustic soda left.

Evaporate the neutral solution until the residue is quite dry. There will probably be some spurting, but this does not matter, as we do not need to weigh the residue in this experiment.

Do you observe the escape of any acid fumes ? Does the solution still remain neutral as it evaporates ?

Do you recognise the residue? Has it the appearance of caustic soda? Dissolve a little in water in a test-tube, and try whether the solution is alkaline or acid. Taste a very small portion. Do you not find it to be *common salt*?

Here then is an unexpected and important result. By the neutralisation of caustic soda with hydrochloric acid we do not obtain merely a mixture of the two with properties which are between those of the acid and the alkali. On the contrary, we find common salt to be produced, a substance which is very different in its properties from either of them. Is a similar result met with when other acids and alkalies are neutralised?

### PROBLEMS

Find the products on neutralising—

1. 20 c.c. caustic potash with nitric acid.
2. 20 c.c. caustic soda with sulphuric acid.
3. 20 c.c. ammonia with hydrochloric acid.
4. 20 c.c. caustic soda with acetic acid.

Evaporate each solution, till ready to crystallise, in a basin or large crucible supported on a metal triangle over a flame. Carefully examine and endeavour to identify the substances which you obtain.

**Salts.**—You have probably found after carrying out these problems that in each case a solid crystalline substance is obtained, which is neutral to litmus and has a salt-like taste. Other neutral mixtures of acids and alkalies might be made, and in each case you would obtain a neutral saline product. These substances are called *Salts*. The term “salts,” then, as we shall use it in chemistry, applies not only to common sea-salt, but also to saltpetre, Glauber’s salt, Epsom salts, sal-ammoniac, and to many others.

So far then as we can say at present a *salt* is a substance, crystalline in appearance, saline in taste, neutral to litmus, which is obtained by the action of an acid upon an alkali.

**Old and New Names.**—You have doubtless noticed that in the previous chapters we have sometimes given two names to the same substance—an old name, such as was used by the alchemists—and a new name adopted in more recent times. Fossils are records of the history of the progress of life on the

earth. Fossil names mark the progress of a science. For example "Spirit of Salt," was the name given quite suitably by the alchemists to the volatile liquid obtained by distilling salt with oil of vitriol. The investigations of chemists during the last 100 years have however led to the abandonment of the name "spirit of salt" and the adoption of the name "hydrochloric acid." And because this latter name is the one that is probably to be found on the bottles in the laboratory, it will be more convenient to use it, although we cannot ourselves as yet attach any meaning to it, as we can do to "spirit of salt." Hence as a matter of convenience we shall adopt in many cases the more recent name, although as a matter of principle it would be better to retain the earlier name until we have found out for ourselves the reasons for employing what is really the more suggestive name.

The older names are still often used in commerce and in the household, while the more recent names find favour in the laboratory.

**Names of Salts.**—Salts are named after the acids from which they are formed. Thus—

Salts formed from sulphuric acid are called sulphates.

„	nitric acid	„	nitrates.
„	hydrochloric acid	„	chlorides.
„	acetic acid	„	acetates.

The salt formed from caustic soda and hydrochloric acid is called *chloride of soda*. *Nitrate of potash* can be made from caustic potash and nitric acid. For reasons which will appear hereafter, chloride of soda is more often called *sodium chloride*, and nitrate of potash *potassium nitrate*.

Some of these substances may be recognised in taste and appearance as already familiar under other names.

It will be convenient to have a list of these names of salts and of some other substances at hand for reference :—

#### TABLE OF SYNONYMS

<i>Old or Household Names</i>	<i>New Names</i>
Oil of vitriol, or vitriolic acid	Sulphuric acid
Spirit of salt, or muriatic acid	Hydrochloric acid
Spirit of nitre, or aqua fortis	Nitric acid



TABLE OF SYNONYMS—*continued*

<i>Old or Household Names</i>	<i>New Names</i>
Vinegar	Acetic acid
Washing soda, or mild alkali	Sodium carbonate
Caustic soda, or caustic alkali	Sodium hydrate
Potash	Potassium carbonate
Caustic potash	Potassium hydrate
Sal-ammoniac	Ammonium chloride
Spirit of sal-ammoniac, hartshorn, or ammonia	Ammonium hydrate
Common salt	Sodium chloride
Nitre or saltpetre	Potassium nitrate
Chile saltpetre	Sodium nitrate
Epsom salts	Magnesium sulphate
Green vitriol	Ferrous sulphate
Blue vitriol	Copper sulphate
Chalk	Calcium carbonate
Quicklime	Calcium oxide
Slaked lime or lime water	Calcium hydrate
Fixed Air	Carbonic acid, carbon dioxide
Fire Air	Oxygen
Vitiated Air	Nitrogen
Inflammable Air	Hydrogen
Nitrous Air	Nitric oxide
Volatile Sulphurous Acid	Sulphur dioxide,

## CHAPTER XIX

### THE DISCOVERY OF "FIXED AIR"

**Is Atmospheric Air the only Kind of Air?**—There are as we well know many kinds of solids, which differ from one another in appearance, in relative density, in melting point, in specific heat, in latent heat of fusion and in many other properties. Nor are we less aware of the existence of different kinds of liquids.

Are there also different kinds of airs? We know that we live in a vast ocean of air, invisible, but around and about us on every hand. Is this atmospheric air the only kind of air? Or are we unaware of other airs, only because they are invisible, formless and vague? In view of the fact that there are so many different solids and liquids, it would be very remarkable if there were but one kind of air. We must at least admit the possibility of the existence of different airs, and it is very desirable to follow up any line of inquiry which promises to lead to the discovery of any air different from that we breathe.

**The Discovery of "Fixed Air."**—As a matter of history, an air differing from atmospheric air was discovered by a Dutch chemist *Van Helmont*, who died in 1644. He gave it the name of *gas*. This word, now so familiar to us, was coined by him for the new air, and is undoubtedly connected with our word *ghost*, and the German word *geist*, a spirit.

His discovery and experiments were, however, forgotten, and it was left for *Black*, Professor of Chemistry at the University of Edinburgh, to establish in 1755 the existence of another air,

differing essentially from atmospheric air. Let us now perform an experiment, which will give us the necessary clue to the discovery of this air.

**EXPT. 1. The Action of Acids upon Chalk.**—Place a little chalk in a test-tube. Pour upon it a few drops of hydrochloric acid. A violent effervescence, probably due to the escape of an air or gas, results. Add more acid. Warm in a flame. Repeat this treatment. There is finally no more effervescence.

**EXPT. 2. To find whether a Gas is evolved by the Action of Acids on Chalk.**—If a gas is really given off, bubbles will probably be seen if it is led through water. Test this point as follows :—

Take a 4 oz flask, and introduce a few grams of chalk (not powdered). Bend a glass tube *a* as represented in Fig. 76. Attach it as a delivery tube to the flask by means of a cork. Let the end dip below the surface of some water in a basin or beaker. Now remove the cork. Pour in about 50 c.c. of weak hydrochloric acid, and at once replace the cork. There is a violent effervescence, and at the same time bubbles are forced up through the water.

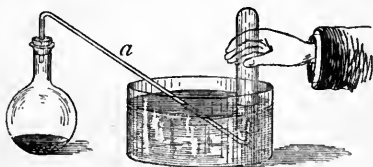


FIG. 76.

Can we collect the gas? Fill a test-tube with water. Close the end with the thumb and invert the test-tube. Place the mouth of the test-tube below the surface of the water in the basin, and then remove the thumb. The test-tube remains filled with water. Move the test-tube so that the end is brought above the end of the delivery tube. Bubbles rise up the test-tube, and displace the water.

There can no longer be any doubt that a gas is evolved by the action of hydrochloric acid on chalk.

**EXPT. 3. Collection of the Gas obtainable by means of Chalk and Hydrochloric Acid.**—If we are to examine this gas we must prepare several samples of it. The apparatus used in the last experiment has the disadvantage that more acid cannot be added without removing the cork, and so admitting

also common air. This can be avoided by passing a funnel through the cork.

(1) Place about 20 grams of chalk (not powdered) in a thick glass bottle. Fit it with a cork, and pass through this a thistle-funnel *a*, and a delivery tube *b*, bent as in Fig. 79. Let the end dip below the surface of some water in a trough or basin.

(2). Fill a glass jar completely with water. Cover the open end with a ground glass plate, and see that no air is included. Holding the glass plate on with the right hand, grasp the jar with the left hand, as in Fig. 77.

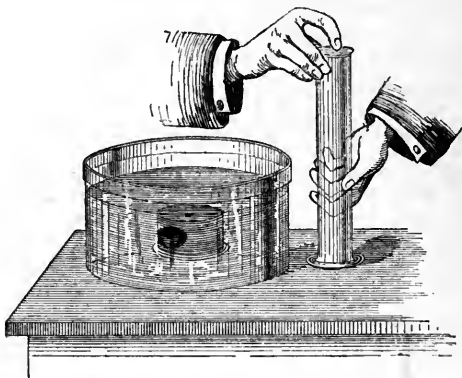


FIG. 77.

(3) Then invert the jar, and place its mouth (still closed by the glass plate) below the surface of the water in the trough, as in Fig. 78.

Now withdraw the plate and allow the jar to stand on the bottom of the trough. Fill three other jars in the same way.

(4) Pour water down the thistle-funnel until the chalk is covered. Then add about 50 c.c. of hydrochloric acid. The commercial quality will do. Bubbles come up through the water. Allow some to escape.

(5) Now place a jar on a little earthenware "beehive-shelf" above the end of the delivery tube, as in Fig. 79. As soon as the water in the jar has been almost entirely displaced, slide it quickly

to one side, slip a glass plate over the mouth (still below the surface of the water), remove the jar, holding the plate in position and place it on the bench. At once place another jar over the delivery tube. Collect in this way five or six samples of the

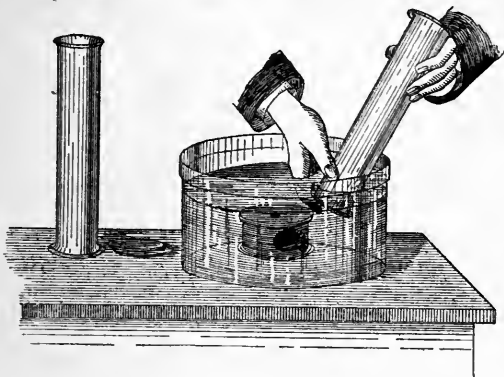


FIG. 78.

gas and leave these standing mouth downwards on the wet glass plates.

This method of collecting the gas is spoken of as *collecting over water*. It is a very convenient method which was devised

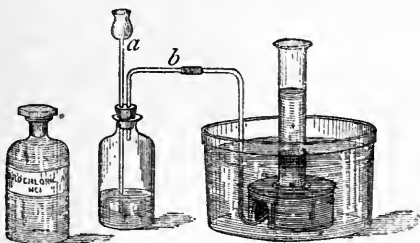


FIG. 79.

by *Priestley*, a chemist, who made great use of it in the latter half of the eighteenth century in experiments to which we shall refer subsequently. The trough or basin used in this

method is often called the *pneumatic trough* (Greek *pneuma*, air or breath).

Black called the air or gas we have obtained from chalk and hydrochloric acid, "*Fixed Air*," because it appeared to be fixed in chalk in the solid state. Black said—"I have given the name of *fixed air*, and perhaps very improperly; but I thought it better to use a word already familiar, than to invent a new name, before we be more fully acquainted with the nature and properties of this substance which will probably be the subject of my further inquiry." The name is so appropriate, as indicating the preparation of the air by means of a *solid* substance and hydrochloric acid, that we shall use it until we find reason for a better name.

But is Fixed Air after all only common air, or is it a different kind of air or gas? Let us try to answer this by experiment. Suppose we try, for instance, whether it has the properties of common air.

### THE PROPERTIES OF FIXED AIR.

**EXPT. 4. To find whether a Taper or Charcoal will burn in Fixed Air.**—(1) Turn one of the jars previously filled with Fixed Air mouth upwards. Remove the plate, and plunge a lighted taper into it. The flame is immediately extinguished, and the burnt part of the wick becomes black instantly. Is the taper injured? Remove it and apply a flame. It is rekindled. Plunge it into the jar again. It is once more extinguished.

(2) Bind some wire round a piece of charcoal. Heat it red hot in a flame, and at once plunge it into a jar of the gas. The glowing charcoal becomes black at once, as though plunged into water.

Here, then, we find a difference between this gas and ordinary atmospheric air.

**EXPT. 5. To find whether Fixed Air is lighter or heavier than Common Air.**—If it is lighter it will probably float above ordinary air; if heavier it will sink. Remove the lid from a jar A of the gas (Fig. 80). After one minute insert a burning taper. It goes out. Now hold the jar upside down for a minute in the position B. Introduce the burning taper again. It still burns. So the gas which put out the taper has fallen out of this jar and is *heavier* than air.

## PROBLEM

To confirm this, try whether you can pour the gas downwards from a full jar into a small beaker full of ordinary air, as in Fig. 81. The gas is invisible, but take care not to spill it, nor to let draughts blow it aside. Can you show the presence of Fixed Air in the beaker?

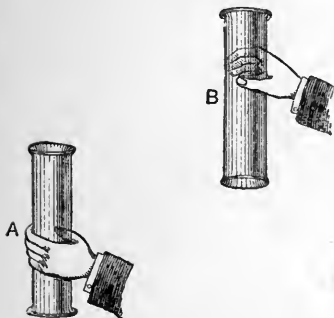


FIG. 80.

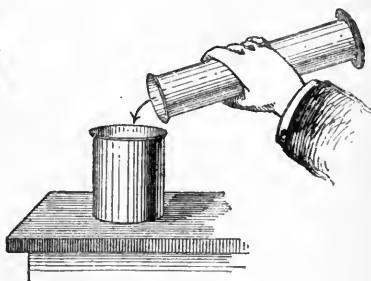


FIG. 81.

The following is another effective way of showing that Fixed Air is heavier than common air.

EXPT. 6.<sup>1</sup> Suspend two thin narrow beakers at the two ends of the arms of a balance. Counterpoise them exactly by adding a little sand to the lighter arm. Take a jar filled with Fixed Air, and holding it in the position shown in Fig. 82, pour its contents into the beaker A. The beaker A descends. Introduce a taper into A and into C, and observe that it is extinguished in A, but burns in C.

EXPT. 7. **Water can be charged with Fixed Air.**—Fill an 8 oz. flask with water and invert it in water. Pass Fixed Air into it until about half of the water is forced out. Then, placing the hand on the upper part of the flask, shake it briskly. In a few minutes the water absorbs the Fixed Air, and taking its place nearly fills the flask as at first. Again half fill the flask with Fixed Air. Shake as before. Is any more of the gas absorbed?

To find whether the properties of water are altered by Fixed Air, run a stream of the gas through water in a beaker. Taste

<sup>1</sup> Lecture Table Experiment.

the water. It has a pleasant faintly acid taste. Try its action on litmus paper. Warm the water. It begins to effervesce.

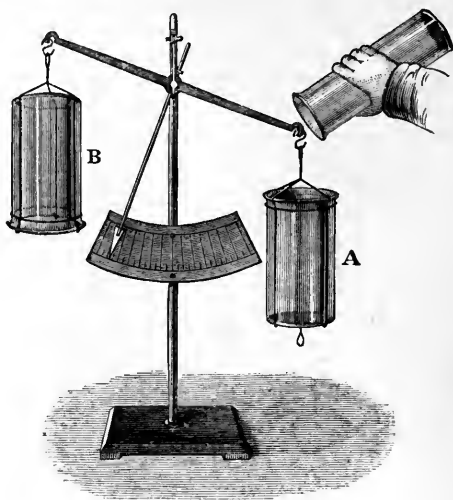


FIG. 82.

## PROBLEM

1. Try the effect of Fixed Air on different liquids. Pass a slow stream of Fixed Air for about half a minute into a few c.c. of several different liquids placed in turn in a test-tube. Take, for instance, hydrochloric acid, caustic soda, lime water, very dilute litmus solution. Find whether there is any *visible* effect in each case.

2. Try the effect of atmospheric air on the same liquids, producing the current of air by means of a foot-bellows, or by an aspirator.

Tabulate your results. Do you notice any difference between the effects of Fixed Air and of common air? Have you found any test which will serve as a ready means of detecting Fixed Air?

Experiment 7 raises the question whether water can dissolve atmospheric air as readily as Fixed Air. Does it dissolve any?

**EXPT. 8. To find whether Water can dissolve any Atmospheric Air.**—Take some water which has recently



been boiled, and has just become *quite cold* again. Half fill a burette with the water and cork the top of the burette. Shake well. Read the level of the water in the burette. Place the lower end of the burette in water so that the level of the water inside and outside is the same. Open the tap. Does any water enter?

Again, have you not noticed the appearance of bubbles on the sides of a beaker or flask in which water is being heated, before the water has begun to boil? These look like air-bubbles, and if they are such, the air must have been dissolved in the water. And if so, the air must be less soluble in hot water than in cold water, while the reverse is almost always the case with solids. This suggests heating a large quantity of water till it boils, in order to see whether any air is expelled.

**EXPT. 9. To find whether Air can be expelled from Water by Heat.**—Fill a tin can completely with water (preferably from a spring. Why?) Fit it with a rubber stopper, and a delivery tube, the end of which is flush with the lower surface of the stopper. Push the stopper well in, so that water is forced into the tube, driving the air out of it. Support the can above a rose burner. Let the other end of the tube dip into a pneumatic trough. Fill a small jar with water and invert it above the mouth of the tube (Fig. 83). Proceed to heat the can until the water boils. Air is carried over into the jar. After boiling a few minutes, remove the burner, and then the delivery tube from the mouth of the jar.

Measure the volume of air collected. It is not nearly so great as the volume of Fixed Air that would dissolve in the same amount of water. But is the air atmospheric air? Close the mouth of the jar, invert it, and introduce a burning taper. It burns just as in ordinary air.

Taste some of the water which has been boiled in the can. It is very flat and insipid. It seems likely that this is due to the loss of the dissolved air. Now drinking water is often obtained at sea from sea water by distillation; but such condensed water is very unpalatable, unless it is afterwards exposed in thin streams to the air. If such water be then heated as in the last experiment, air is given off again.

Moreover, since atmospheric air is necessary for the life of man, it seems probable that it is the air dissolved in water

which supports the life of fishes and all aquatic animals. This is confirmed by the fact that no fish can live in water which has been boiled.

We may, therefore, conclude that the small amount of air which is found dissolved in water is really obtained from the atmosphere.

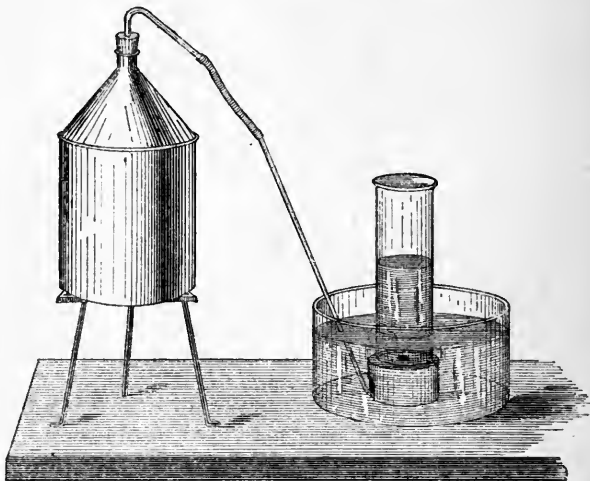


FIG. 83.

#### PROBLEM

Find whether Fixed Air is identical with the gas contained in soda-water.

**The Source of the Fixed Air.**—Observe that while we have proved that chalk in contact with hydrochloric acid gives off a gas which is different from atmospheric air, yet we have no evidence whatever to lead us to conclude whether the Fixed Air comes entirely from the chalk, or entirely from the acid, or partly from one and partly from the other, or from both. We cannot at present say which is the case.

But is chalk the only solid which, in contact with hydrochloric

acid, will evolve Fixed Air? And is hydrochloric acid the only acid which is capable of doing this?

### PROBLEMS

(1) Find whether marble, limestone, calc spar, sand, plaster of Paris, whitening, oyster shells, egg shells, washing soda, do or do not give off a gas in contact with hydrochloric acid. If a gas is given off, find in each case whether it is Fixed Air.

(2) Find whether chalk in contact with sulphuric acid, nitric acid or acetic acid gives off Fixed Air.

**Laboratory Method of preparing Fixed Air.**—Knowing now that Fixed Air is heavier than common air, we can devise a readier way of collecting it than that described in Experiment 3, which is also open to objection because some of the gas disappears into the water.

**EXPT. 10. Collection of Fixed Air by Downward Displacement.**—Attach to a bottle in which Fixed Air is to be produced a glass tube bent at right angles, so that it may reach to the bottom of a jar (Fig. 84). Place about 20 grams of marble in small pieces (not powdered) in the bottle. Add water first, and then an equal volume of hydrochloric acid. The Fixed Air which is evolved will collect at the bottom of the jar and gradually displace all the air. The jar is known to be full of the Fixed Air when a taper held just below its mouth is extinguished.

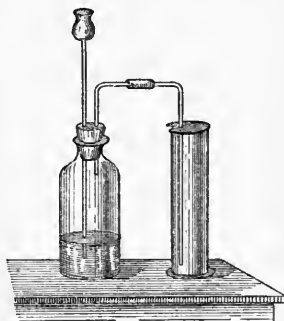


FIG. 84.

This apparatus is to be used in future when Fixed Air has to be prepared.

**Summary.**—The main result, then, at which we have arrived is that there are several substances which in contact with acids give off a gas which is essentially different from common air, for a taper cannot burn in it; it is heavier than common air; it has an acid character; it renders lime-water

milky, and water will dissolve more than an equal volume of it. This air we have named "Fixed Air."

Black did not leave the matter here, but measured the amount of Fixed Air which could be obtained by means of a given weight of chalk, and subjected, chalk itself to a thorough examination.

We shall, for the present, however, leave the study of chalk, in order to follow up other lines of inquiry, which were successfully pursued by a few chemists in the seventeenth and eighteenth centuries. Subsequently, when our hands have greater skill, and our faculties are better trained, we shall return to Black's famous research, which is a model of what an investigation should be.

## CHAPTER XX

### THE STUDY OF FIRE AND AIR

**THE PHLOGISTIC THEORY OF FIRE.**—We have all watched a bright fire, and delighted in the long tongues of flame, flickering and darting upwards, and we know that after it has died out dead ashes alone remain. What has occurred? Is the flame simply something escaping from the coal? Men used to think so. Or again, take the following instance of another kind of flame.

**EXPT. I. The Burning of Magnesium.**—Take a strip of magnesium ribbon, 10 cm. in length, and holding one end in a pair of tongs, kindle the other end. It burns brilliantly, and a white ash or calx is left. Is this dazzling light the evidence of something parting from the magnesium and leaving the calx?

This idea, the first perhaps that is likely to occur to any one, agrees with one which was put forward in the seventeenth century by Becker, a distinguished chemist, and by Stahl, a physician to the King of Prussia. They taught that all combustible bodies were compounds containing at least two constituents, and that when combustible bodies were burnt, one of these constituents escaped while the other remained.

In the burning of the magnesium, for instance, they would have said that a combustible constituent escaped which had been contained in the metal. They gave the name “phlogiston” to this combustible constituent, and their Theory of Combustion was known as the *Phlogistic Theory*.<sup>1</sup>

An advantage of this theory was that all cases of combustion

They said Phlogiston was *materia aut principium ignis, non ipse ignis*—the substance or principle of fire, not fire itself.

were looked at from the same standpoint. For it was considered that all combustible substances contained the same phlogiston, and in the burning of a candle, an oil-lamp, a fire, sulphur, or a metal such as magnesium, they saw alike the escape of this phlogiston.

Now these chemists paid no attention in their investigations to the question whether or no the air had anything to do with fire and flame. This was never considered. Nor were they accustomed to weigh the materials with which they began an experiment, nor to determine the weight of anything left at the end. In fact, measurement was entirely neglected by them, just as it had been by the searchers after the philosopher's stone.

**THE STUDY OF FIRE IN THE SEVENTEENTH CENTURY.**—A result of the neglect of measurement on the part of the early chemists, was that although Fire had been an object of wonder and attention since the earliest ages, it was not until the middle of the seventeenth century that any real knowledge of the phenomenon was obtained. This we owe to a little band of chemists in Oxford.<sup>1</sup> It will now be our object, following their guidance, to begin for ourselves the study of Fire and Flame, so as to learn something of what really does occur when any substance is burnt in the air.

*Robert Hooke*, one of these Oxford chemists, gave in 1665 the following description of the method of converting wood into charcoal, which can readily be repeated, and which will give us an important clue.

**EXPT. 2.<sup>2</sup> Air is Necessary for the Burning of Charcoal.**—"The body to be charr'd or coal'd may be put into a Crucible, Pot, or any other Vessel, that will endure to be made red hot in the Fire without breaking, and then cover'd over with Sand, so as no part of it be suffer'd to be open to the Air, then set into a good Fire and there kept till the sand has continu'd hot for a quarter, half, an hour or two, or more, according to the nature and bigness of the body to be coal'd or charr'd, then taking it out of the Fire, and letting it stand till it be quite cold,

<sup>1</sup> See Prof. H. B. Dixon's address before Section B of the British Association, Oxford, 1894.

<sup>2</sup> Lecture Table Experiment.—A piece of wooden penholder, buried in sand within a platinum crucible, may be heated with a Bunsen burner.

the body may be taken out of the Sand, well charr'd and cleans'd of its waterish parts."<sup>1</sup>

If, however, whilst the crucible is still hot the contents are turned out upon an iron plate so that the body is exposed to the air, Hooke observes that "when it comes into the free air, it will take fire, and readily burn away."

This experiment is very important, for it seems to show that the burning of the charcoal is dependent upon the presence of air; indeed, this is the conclusion which Hooke drew, for he said:—

"From the experiment of charring of Coals (whereby we see that notwithstanding the great heat and the duration of it, the solid parts of the Wood remain, whilst they are preserv'd from the free access of the air undissipated) we may learn, that which has not, that I know of, been publish'd or hinted, nay, not so much as thought of, by any; and that in short, is this—that the Air in which we live, move and breathe, and which encompasses very many, and cherishes most bodies it encompasses, is the universal dissolvent of all Sulphureous bodies" (*i.e.*, of all combustible bodies).

Now if it be true that the air is necessary for the combustion of any body, it follows that a substance, for example sulphur, will not burn in absence of air. This *Robert Boyle* proved in 1660 to be the case by an experiment which it is not easy to repeat. He lowered some sulphur on to a hot iron plate in a vessel from which the air had been pumped, and he found that it would not burn, but on allowing a little air to enter, "divers little flashes could be seen," which were extinguished on exhausting the air again.

## THE HEATING OF METALS IN THE AIR

In the course of the seventeenth century a number of chemists directed their attention to the effect of heat upon metals, and very important results were obtained from their observations.

**EXPT. 3. To observe the Effect of Heating Lead in the Air.**—Put some lead in an iron spoon.<sup>2</sup> Heat it over the Bunsen flame. After melting, an earthy-looking scum or calx

<sup>1</sup> Hooke, *Micrographia*. See Alembic Club Reprint, No. 5, p. 42.

<sup>2</sup> In the absence of sufficient spoons, crucible lids supported on home-made pipe-clay triangles can be used.

forms upon the surface of the molten metal. Scrape the calx to one side with an iron wire. A mirror-like metallic surface is exposed, which quickly becomes coated over again.

In what way may this calx have arisen? Two causes readily suggest themselves. Firstly, it is possible that something has been driven out of the metal by the heat, leaving the calx as a residue. This is the view which was taken by many chemists 200 years ago.<sup>1</sup> Secondly, it is possible that under the influence of the heat some portion of the air becomes attached to the metal, whereby the calx is formed. Now it is evident that we can decide between these two cases by means of the balance. For if something has been driven out of the metal there must be a *loss* of weight, while if any has been added to it, there must be a *gain* in weight. Let us, therefore, put the matter to the test, and weigh the metal before and after it is heated.

**EXPT. 4. To find whether Lead gains Weight when Heated in the Air.**—Weigh a half ounce crucible. Weigh

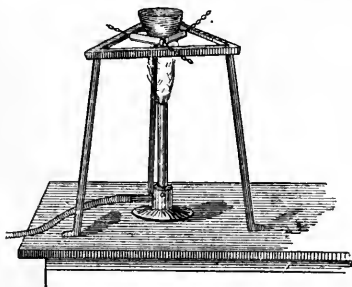
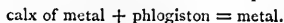


FIG. 85.

about 3 grams of clean lead parings into it. Support the crucible on a pipe-clay triangle upon a tripod stand (Fig. 85). Heat it (uncovered) for about twenty minutes over a Bunsen flame. It is well to scrape the calx to one side from time to time with an iron wire, so as to expose a fresh surface, taking care, however, that none is removed from the

crucible. Allow it to cool. Weigh the whole again. Is there any increase in weight? Repeat for the same length of time. Is there any further increase in weight? If the calx formed on the surface of the lead is red, it is called *red-lead*. If yellow, it is called *litharge*.

<sup>1</sup> This view was taken by the Phlogistic chemists. They believed the heat drove a fiery material *phlogiston* from the metal, and therefore they regarded the calx as a simpler substance than the metal. Their view can be expressed thus:—





It is so important to decide whether there is an increase in weight when a metal is heated in the air that experiments with other metals must be made.

**EXPT. 5. To find whether Magnesium gains Weight when heated in the Air.**—Weigh a half ounce crucible with its lid. Weigh out 0.2 to 0.3 gram of narrow magnesium ribbon or wire on to the pan bearing the crucible. Roll the ribbon up into small compass, and place it in the crucible. Support the crucible with the lid upon a pipe-clay triangle on a tripod stand. Then apply a flame gradually until the bottom of the crucible is red hot. This should take five minutes. Continue heating strongly for another five minutes, but lift the lid about every half minute *very slightly* so as to let in air but not to allow the escape of fumes. Allow to cool. Weigh again. Is there an increase in weight?

**EXPT. 6.<sup>1</sup> To find whether Iron gains Weight when heated in the Air.**—Suspend a horse-shoe magnet from the hook of a balance pan as shown in Fig. 86. Place a round of filter-paper on the pan below. On the paper place a little heap of finely divided iron.<sup>2</sup> Lift the paper and allow the magnet to pick up as much iron as possible. Weigh together the magnet, iron and paper. Let the flame of a Bunsen burner play gently against the iron on the magnet for a few seconds. The iron begins to glow. Remove the flame. The glow continues and spreads. The iron is burning. Let it cool. Weigh again. It is heavier.

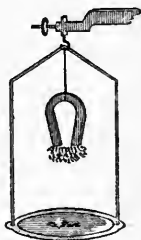


FIG. 86.

In the seventeenth century experiments were frequently made upon the calcination of *antimony* by means of the sun's rays concentrated upon it by a powerful burning glass. A small cone of the metal was weighed and placed upon a slab of marble and the point of the refracted rays directed on to the point of the cone. The metal immediately fumed and was converted into a white calx, which was separated with a knife, and the rays directed on to the rest till the

<sup>1</sup> Lecture Table Experiment.

<sup>2</sup> *Note for Teacher.*—Iron filings are not so suitable as freshly prepared *ferrum redactum*. If spoiled or rusty, heat the iron in a hard glass tube in a current of coal gas over a Ramsay burner. Wait until quite cool before removing it.

whole became white. It was then weighed again. As a writer said in 1630 :—

“It is a remarkable thing that although in this calcination the antimony has lost much of its substance, by the vapours and fumes which are copiously exhaled, yet its weight augments instead of diminishing.”

From what source comes this increase in weight, when a metal is calcined? In view of all the experiments we have made it is impossible to conclude that the increase arises from anything else but the *fixing of particles of air in the metal during the heating*. We can scarcely believe that the increase is simply due to heat itself having weight, because Experiment 2 clearly showed that the presence of air was necessary for the burning of charcoal.

**THE DISCOVERY OF “FIRE-AIR.”**—The point at which we have arrived is that air is necessary for the combustion of charcoal, and that when a metal is heated it increases in weight, probably because it combines with particles of air. Suppose then we heat a substance in air in a *closed* vessel, what is likely to occur? If our conclusions are correct, is it not probable that some of the air will disappear?

It is not easy to test this in the case of metals. But in the year 1674 a very remarkable substance called *phosphorus* was discovered, which is particularly suitable for experiments upon this subject.

Phosphorus can be prepared in two very different forms ; one is a yellow waxy substance, the other a dark red powder.

It was first prepared from urine by an alchemist named Brand in Hamburg by a secret process. A few years afterwards it was discovered how to prepare it from bone ashes. Even at the present time the details of its preparation are kept secret by the few firms which are concerned with its manufacture.

Its properties are so remarkable that it was exhibited as one of the wonders of nature to King Charles II.

**EXPT. 7. To observe the Burning of Phosphorus in the Air.**—Obtain a gas jar. Cut a clean piece of phosphorus, about the size of a pea, and dry it with blotting-paper. Place it in a “deflagrating spoon.” Warm one end of a glass rod in a flame, and then touch the phosphorus with it. The phosphorus catches fire. Introduce it into the jar. Dense white fumes are

evolved, and settle on the inside of the jar, and presently the flame dies out. Remove the spoon. Some phosphorus remains unburnt, so perhaps it can be re-kindled. Try. Yes, it will burn further in the air outside the jar.

How has the white powder arisen? Has a fiery material been driven out of the phosphorus leaving the powder as a residue? Or has some portion of the air become attached to the phosphorus, so forming the white powder?

We can much more readily try whether any air disappears when phosphorus is burnt than when a metal is burnt or heated in the following way.

**EXPT. 8. One-fifth of the Air in a Closed Flask disappears when Phosphorus is burnt in it.**—Take a clean and dry 8 oz. flask.<sup>1</sup> Weigh it. Select a tightly fitting rubber stopper.<sup>2</sup> Cut a clean piece of phosphorus,<sup>3</sup> about the size of a pea, and dry it with blotting-paper. Place it in the flask and close it firmly with the stopper. Weigh the whole. Then, holding the flask by the neck and away from your face, warm it everywhere except where the phosphorus is. Turn the flask and let the phosphorus fall on to its heated sides. It burns in the enclosed space (probably without bursting the flask, but beware!). Replace the flask on the pan of the balance. It may appear lighter than before. (Why?) Allow it to cool. Its weight has not been changed, showing that no material substance, at any rate, has passed through the walls of the flask. Now place the neck of the flask below the surface of some water, and remove the stopper. Observe that the water rises into the flask; shake the flask until the white deposit on the sides of the flask and the fumes are dissolved. Replace the stopper. Remove the flask from the water and dry the outer surface. Weigh it, and so find how much water has entered the flask.<sup>4</sup>

<sup>1</sup> A flask can be dried by waving it over the flame of a Bunsen burner and sucking out the air by a glass tube placed in the flask. Allow it to cool completely before weighing it.

<sup>2</sup> A cork soaked in melted paraffin may be used.

<sup>3</sup> *Caution.*—Phosphorus is *kept* under water, and *cut* under water. It is handled with tongs. Before burning it is dried by pressing between folds of filter-paper. It catches fire of itself, at a temperature below that of boiling water. But it cannot burn in absence of air. If the hands are burnt by phosphorus, cover the wounds with a mixture of olive oil and slaked lime, called "carron oil."

<sup>4</sup> You may more rapidly, but not so accurately, measure the water which has entered the flask by pouring it into a graduated cylinder. Measure in the same way the whole volume of the flask.

Find also how much it can contain when filled up to the stopper. It will be found that the water which enters the flask is just about one-fifth of that which it can contain.<sup>1</sup>

It is evident, therefore, that in the combustion of the phosphorus, one-fifth of the volume of the air disappears. What has become of it? It cannot have escaped from the flask, for there was no loss of weight. Is it not likely that it has united with the phosphorus, forming these white fumes which filled the flask, and which dissolved in the water? (Test the water with litmus paper and observe that it is acid.)

"But," as Lavoisier, a famous French chemist, said when he was studying the matter, "there is a great difference between conjecture and proof, and it is necessary that it should first be established that a combination of any kind is formed with phosphorus during its combustion."

Now, if the air which has disappeared has really united with the burning phosphorus, the phosphorus must have increased in weight, just as the metals did. In order to make sure whether this is so or not, the burning of phosphorus may be carried out as follows :—

**EXPT. 9.<sup>2</sup> To find whether the Products of Burning Phosphorus are heavier than the Original Phosphorus.**—Take a piece of wide glass tubing about 6 inches long, drawn out at one end (Fig. 87). Place a small piece of phosphorus at D. Pack about 3 inches of it from B to C with fine asbestos. Close the end with a stopper through which a glass tube passes. Test by sucking to be sure that air can filter through the asbestos. Weigh the whole. Then fix the tube in a horizontal position, and connect the end B with an aspirator. Ignite the phosphorus, and the fumes will be carried towards B and caught by the asbestos. Finally, when the phosphorus is consumed, weigh the tube again. There is a very decided increase in its weight.

There can therefore be no doubt that when phosphorus is burnt in a definite volume of air, one-fifth of the air disappears, and it does so because it unites with the phosphorus forming a white solid substance.

<sup>1</sup> The experiment in this form is due to the Swedish chemist Scheele (about 1773 ; see Alembic Club Reprint, No. 8, p. 13).

<sup>2</sup> This experiment was suggested by Prof. H. E. Armstrong.

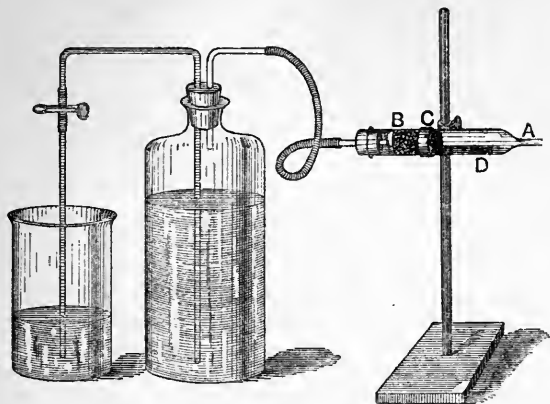


FIG. 87.

Let us now make two simple experiments with a *candle*.

**EXPT. 10. A Burning Candle needs Fresh Air.**—Attach a candle to a deflagrating spoon. Light it. Put it inside a tall dry glass jar, closing the top of the jar with the plate attached to the spoon (Fig. 88). The flame burns for a time, sinks and goes out, while the inside of the jar becomes dim. Has anything hurt the candle? Take it out, leaving the jar covered with a glass plate. Try whether the candle can be lighted again. It burns: clearly it is in no way spoilt. Has the air been spoilt? Lower the burning candle into the jar again. It goes out almost immediately. Then it is the air which has been altered by the burning candle.

**EXPT. 11. To find whether Air disappears when a Candle is burnt in a Closed Vessel.**—The following is an experiment which Mayow showed in Oxford so long ago as 1670. He wrote:—

“Let a lighted candle be so placed in water that the burning wick shall rise about six fingers’ breadth above the water; then

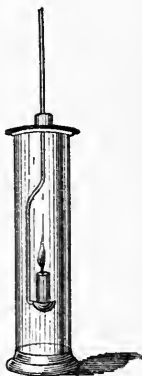


FIG. 88.

let a glass vessel <sup>1</sup> of sufficient height be inverted over the candle. Care must be taken that the surface of the water within the glass shall be equal in height to that without, which may be done by including one leg of a bent syphon within the vessel while the other opens outside. The object of the syphon is that the air enclosed by the vessel and compressed by its immersion into the water may escape through the hollow syphon. When the air ceases to issue the syphon is immediately withdrawn so that no air can afterwards get into the glass. In a short time you will see the water gradually *rising* into the vessel while the candle still burns.”

From all that has been brought forward it is now evident that there can be no combustion without the presence of air, and that when a substance is burnt it combines with some of the air, and on that account increases in weight. It has been noticed, however, that when a substance is burnt in a closed volume of air only one-fifth of the air disappears and not the whole. This fact makes it almost certain that the air is not really a single substance, but contains at any rate two gases, one of which alone (occupying one-fifth of the whole) is concerned in combustion; this hypothetical portion of the air may conveniently be called *Fire-Air*.

Would another possible explanation of the disappearance of one-fifth of the air be that the flame of phosphorus or of a candle causes air to contract?

**THE DISCOVERY OF “VITIATED AIR.”**—Let us now remove the Fire-Air from a given volume of common air, by burning some phosphorus in it, and try what kind of air is left.

**EXPT. 12. To remove the Fire-Air from Common Air, and to find whether the Air which is left supports Combustion.**—Take a large glass bell-jar A (Fig. 89) with a stopper at the neck, and graduate it into six equal divisions.<sup>2</sup> Float a small porcelain dish D containing a piece of dry yellow phosphorus in some water in a dish B. Cover it with the bell-jar, so that the water rises to the fifth mark. Touch the phosphorus

<sup>1</sup> A flask with a short neck is suitable.

<sup>2</sup> The experiment can be made with an ordinary bell-jar. In that case some of the gas must be decanted into gas jars for testing.

with a hot glass rod, and quickly replace the stopper. Dense fumes are given off, which soon dissolve in the water, and the water finally rises almost to the fourth mark, if water is poured into the dish so that its level inside and outside the jar is the same. Remove the stopper. Introduce for a moment a lighted taper. It is at once extinguished. Place a small piece of phosphorus or sulphur in the cup of a deflagrating spoon. Ignite it. Introduce it into the jar for a moment. It is extinguished. It was also found more than 100 years ago that animals died almost immediately on being introduced into this portion of common air which was left after the removal of Fire-Air. Let us for the present call this air *Vitiated Air*.

Now it is one of the most characteristic of the properties of common air that it is necessary for the life of animals. This has long been known, and it was clearly proved by Boyle in experiments which the student is recommended not to repeat; he placed cats and mice, sparrows, fishes, and tadpoles in the receiver of his air-pump, and terminated their lives by pumping out the air. It has been stated above that when the Fire-Air has been removed from common air animals cannot live in the air which remains.

The conclusion which necessarily follows is that it is only that portion of the air which we have called Fire-Air which enables animals to breathe and is concerned in what is called respiration.

Is this Vitiated Air, however, the same as Fixed Air?

**EXPT. 13. To find whether Vitiated Air is the same as Fixed Air.**—Transfer some of the Vitiated Air left in the jar in Experiment 12 to another jar. Pour in some clear lime-water, and shake well. No turbidity results. Then Vitiated Air must be a different kind of air from Fixed Air, though both of them extinguish a lighted taper.

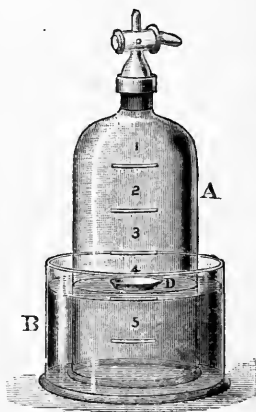


FIG. 89.

**Common Air is not a Simple Substance.**—There can therefore be no doubt that ordinary air contains at any rate *two* airs. One we have called Fire-Air ; this it is which enables combustible substances to burn and animals to breathe. The other is entirely different from Fire-Air, for it can neither support combustion nor life.

This extremely important conclusion, namely, that common air contains two entirely different airs or gases, was first arrived at about 1670, from experiments similar to those we have seen, carried out by the Oxford chemist *Mayow*. This discovery, however, attracted little attention, and was soon entirely forgotten. It was left for the Swedish chemist *Scheele* to remake it about the year 1772, and he it was who gave the two airs the names *Fire-Air* and *Vitiated Air*.

**The Nature of Flame and Fire.**—Can we now understand the nature of the flame of phosphorus when it is burnt in a closed vessel ? Nothing material escapes from the vessel. The original contents were Fire-Air and Vitiated Air. During the experiment light and heat are given out. After the experiment the white dust on the glass is the Fire-Air combined with the phosphorus. As Hooke said of an experiment in 1665 :—

“It seems reasonable to think that that shining transient body which we call *Flame* is nothing else but a mixture of Air and volatil parts of combustible bodies which are acting upon each other whilst they ascend, and that *this action produces a very great heat*, and that which we call Fire ; and that this action is perform’d with so great a violence, and does so rapidly agitate the smallest parts of the combustible matter that it produces in the medium of the Air the action or pulse of Light.”

This theory of flame applies equally to the burning of sulphur or of a candle in the air. The heat and light of a burning candle, for instance, we are to regard as the *result* of the violence of the action of the Fire-Air in the atmosphere upon the combustible material of the candle, the heat being so great that the burning materials are heated white hot, and give out a bright light. The flame of a candle, therefore, is the visible evidence of action going on between the Fire-Air and the combustible matter of the candle.

It is this vigorous action between Fire-Air and combustible material, resulting in heat and light, to which the name *com-*



*bustion* is commonly given, though we shall find afterwards that the word must be used in a somewhat wider sense.

### EXERCISES AND PROBLEMS

#### Fire and Air

1. Observe the effect of heating the following substances on porcelain lids :—

Zinc.

Copper.

Tin.

Magnesium.

Antimony.

Charcoal.

Iron.

Sulphur.

Solder.

2. Find whether there is any change in weight when the following metals are heated in open porcelain crucibles :—

Solder.

Tin.

Zinc.

3. Find whether any air disappears when some sulphur, melted and ignited in an iron spoon, is poured into a dry flask, which is then immediately stoppered.

4. Burn a piece of phosphorus in a closed flask as in Experiment 8, p. 167. Weigh the flask. Then remove the stopper, replace it, and weigh again. Is there any change in weight?

## CHAPTER XXI

### THE RUSTING OF IRON

**THE RUSTING OF IRON.**—In the course of our study of Fire and Air we found that finely divided iron would actually burn in the air, and that in burning it gained in weight. Iron, however, often undergoes some kind of change even at the ordinary temperature; it readily rusts. The bright steel of a bicycle which has been ridden in wet weather is quickly corroded. The rails of a disused railway rapidly acquire a thick coating of red rust.

Take any piece of old iron which has stood in the open air for some time. Scrape off some of the rust and examine it carefully. Have you met with anything like it before in your experiments?

Here then is a problem ready to our hand. What is it that occurs in such cases to iron or steel? We cannot answer this all at once.

The balance, however, will tell us whether iron in rusting loses or gains anything, and secondly it will perhaps be well to try whether any air is set free or disappears in the process.<sup>1</sup>

**EXPT. 1. To find whether Iron changes in Weight when it Rusts.**—Weigh a watch-glass. Place upon it 3 or 4 grams of bright iron filings, and a long iron nail. Weigh

<sup>1</sup> The teacher will probably find it advisable to carry out himself most of the experiments in this research before the class. They require time, and may suitably be continued while the next research on the Discovery of Oxygen is being pursued. For instance, Experiments 1 and 2 may be started together in the course (perhaps best at the end) of one day, and the results examined during the next.

again. Moisten the filings thoroughly with a little water, and stir thoroughly, carefully avoiding the loss of any filings. Place the watch-glass in a warm place near a fire or stove. Cover it with an inverted beaker or shade. Allow it to stand for one or two days. Then place the glass upon a sand-bath, and warm it until quite dry. Weigh. There is a gain in weight. This must have arisen either from the air or from the water or from both.

**EXPT. 2. To find whether some of the Air disappears when Iron Rusts in it.**—Cut a cork so that it can be pushed up a glass jar and fit air-tight. Bend a pin or a small piece of wire into a hook and fix it to the cork. Make a small bag of fine muslin. Place some bright iron filings in it and suspend it from the hook. Push the cork right up to the closed end of the jar. Place the jar mouth downwards in a basin or beaker containing a little water (Fig. 90). Put the whole aside in a warm place until next lesson. After a day or two the level of the water is found to have risen in the jar. Pour water into the beaker so that the level of the water is the same inside and outside the jar. (Why is this done?) Mark the level of the water in the jar. Leave it for another day and then mark the level again. Repeat this. At last the level of the water remains fixed. Find what fraction of the air in the jar has disappeared by measuring with a rule the length of the jar and the length of the column of air which is left. Leave the jar as it is for Experiment 5.

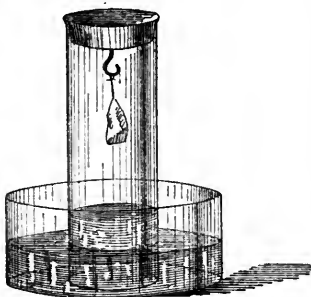


FIG. 90.

¶ We have now proved that when iron rusts in the presence of air and water :—

1. The iron gains in weight.
2. About one-fifth of the air disappears when the rusting occurs in a closed volume of air.

It seems to be certain, therefore, that the increase in weight

of the iron is due (in part at any rate) to a portion of the air becoming fixed in or combined with the iron. It does not at present seem clear from these experiments whether the water has any share in the rusting. Common experience has certainly led people to imagine that water is more than anything else the cause of rusting. Popular opinion, however, is perhaps at fault, and must not be too readily trusted.

Will it not be well to try next whether (1) water in the absence of air, and (2) air in the absence of water can cause iron to rust?

**EXPT. 3. To find whether Iron rusts in Water free from Air.**<sup>1</sup>—Fill a small flask with water. Heat it to boiling

for at least five minutes, so as to expel all dissolved air. Meantime rub some iron nails bright with emery paper. Place them in the boiled water, and close the flask with a tightly-fitting rubber stopper. Set it aside till next lesson. No rusting occurs even after many days.

**EXPT. 4. To find whether Iron rusts in Air free from Moisture.**

—Take a dry glass jar. Suspend a bag containing bright iron nails or filings within

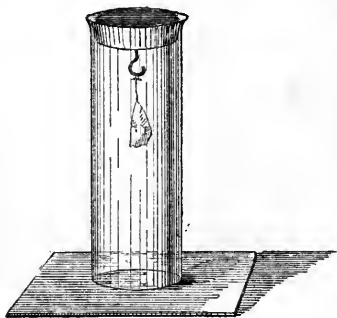


FIG. 91.

the jar exactly as in Experiment 2. Place the jar mouth downwards upon a well-greased plate, and set it aside in a warm place till next lesson (Fig. 91). Then try whether any of the air has disappeared by removing the plate from the mouth of the jar under water. The water does not rise in the jar. Moreover, the iron remains bright.<sup>2</sup>

Experiments 3 and 4, therefore, satisfy us that neither water alone nor air alone can be the cause of the rusting of iron. Consequently, it is clear that iron only rusts in air when the air is moist. The formation of rust on iron is then due to two

<sup>1</sup> Experiments 3 and 4 should be started together at the end of a lesson.

<sup>2</sup> Ordinary air always contains a little invisible aqueous vapour, and consequently the iron may rust to a very small extent

actions : firstly, the action of the air ; and secondly, the action of moisture. Although in the present instance we shall only consider the first action in detail, the student should always accustom himself in studying an action to observe so far as possible all that goes on and all the conditions necessary for its success.

Now when iron was allowed to rust in moist air in Experiment 2, some of the air, but not the whole, disappeared and combined with the iron. Moreover, the fraction of the air which disappeared was *one fifth*, which was exactly the same fraction of the air that vanished when phosphorus was *burnt* in the air (*see* p. 167, Experiment 8). Is it the fact then that that portion of the air which disappeared in Experiment 2 was simply the Fire-Air which it contained? If so, the air which remains must be Vitiated Air. We can put this to the test.

**EXPT. 5. To find whether the Air left after the Rusting of Iron in it is Vitiated Air.**—Fill the beaker containing the jar in Experiment 2 up to the brim. Gradually raise the jar until you can push a well-greased plate underneath it. Now invert the jar on the table with the plate uppermost. Light a taper. Also place a little red phosphorus in a deflagrating spoon, and kindle it. Remove the plate, and at once plunge the burning phosphorus within the jar. It is extinguished. Introduce quickly the burning taper. It also is put out.

We can safely conclude, therefore, that the air which remained after the rusting was Vitiated Air.

**The Rusting of Iron is a Sort of Slow Burning.**—We have therefore proved beyond doubt that in the rusting of iron in a closed volume of moist air the observed gain in weight is due, in part at any rate, to Fire-Air combining with the metal. Now when iron burns in air the iron also combines with the Fire-Air ; but in this case the action is completed in a few seconds, while the rusting of iron takes place very slowly. It seems therefore as though in the rusting of iron there is really a very slow action between Fire-Air and iron, so slow that no heat nor light that we can detect is produced. It seems, indeed, to be a case of very slow combustion (*see* p. 172).

**THE SMOULDERING OF PHOSPHORUS.**—The above interesting result leads us to ask whether any other sub-

stance can combine without flame in this slow and gradual way with Fire-Air. Phosphorus is a likely substance to try for this slow combustion, since it begins to smoulder and fume as soon as it is exposed to the air.

**EXPT. 6. To observe the Smouldering of Phosphorus.**—Remove a small piece of yellow phosphorus from the bottle. Dry it quickly by pressing it gently between folds of filter-paper. Place it upon a porcelain crucible lid. A white smoke begins to arise from the phosphorus. Sometimes the phosphorus will soften and melt, showing that it is warm, and occasionally it will even become hot enough to take fire by itself.

Take the phosphorus to a dark place. It is seen to glow. Hence its name *phos-phorus*, the light-bringer, from the Greek words *phos*, light, and *phero*, I bring.

**EXPT. 7. To find whether some of the Air disappears when Phosphorus is left exposed in it.**—Take a stout piece of copper wire about 10 inches long. Bend one end into a circle, so that the rest of the wire can stand upright upon this as a support. Place a stick of phosphorus in a porcelain basin containing water. Warm the basin until the phosphorus is melted. Then attach a rubber tube to a glass tube, and cautiously suck some of the molten phosphorus up into the glass tube. When the phosphorus has solidified, remove it from the tube, and bind it under water to the straight end of the copper wire by means of some fine iron or copper wire. Place the wire bearing the stick of phosphorus in a glass dish (Fig. 92). Invert a small jar over the phosphorus, and support it by a clamp or by the ring of a retort-stand. Then pour water into the dish to a depth of 1 or 2 inches. Set the whole aside until next lesson.<sup>1</sup>

After two or three days the water is found to have risen inside the jar. Pour water into the dish so that its level is the same inside and outside. Estimate what fraction of the air has disappeared. Is it not about one fifth? Test the water with litmus.

It is probable then, that the air remaining is Vitiated Air. We can readily put this to the test.

**EXPT. 8. To find whether Vitiated Air is left after**

<sup>1</sup> The phosphorus may be exposed in air over mercury with advantage. The resulting fumes are then more visible.

**the Exposure of Phosphorus in Ordinary Air.**—Light a taper, and try whether it will burn in the air left in the jar, used in Experiment 8. It is extinguished at once. Replace the stopper.

**The Smouldering of Phosphorus is a Sort of Slow Burning.**—We have now established the following points with regard to that which occurs when phosphorus is exposed in a closed volume of air.

1. One fifth of the air disappears.
2. The air which is left has the properties of Vitiated Air.

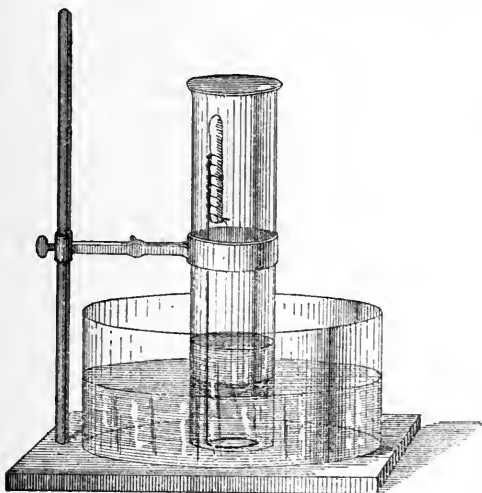


FIG. 92.

We have therefore strong reasons to conclude that phosphorus combines even at the ordinary temperature with Fire-Air. The action appears to be precisely the same in kind as that which occurs when phosphorus is burnt in air, and the difference between the two cases seems to be simply one of rapidity and intensity.

Phosphorus like iron, then, can combine with Fire-Air either very slowly or extremely rapidly. In the one case no action is visible, in the other much heat and a brilliant light are produced. It seems probable, indeed, that the heat and light in the second

case are the *result* of the rapidity and violence of the action between the Fire-Air and the phosphorus or iron. This, it will be remembered, was the conclusion we came to at the end of our last research (*see* p. 172).

When the combination of phosphorus with Fire-Air is extended over hours or days, it seems probable that the action is scarcely visible because it is so slow. A faint glow can, however, be seen if the experiment is made in a dark room.

We have, therefore, confirmed by this study of the slow changes which phosphorus and iron undergo in the air the conclusion we had already come to about Flame and Fire. But this was not at all the actual object we had in view when we began to study the rusting of iron. It is not, however, any less valuable on that account, and as a matter of fact it frequently happens that in the pursuit of knowledge we set out with some one object in view, and we arrive at an entirely different and unexpected discovery, yet perhaps even more important. But if our researches are to lead us thus into wider knowledge, we must be careful to keep our minds open, quick to observe any new fact, and willing to endeavour to understand it.

## PROBLEMS

### The Rusting of Iron

1. Observe the effect of heating iron rust in a test-tube.
2. Find whether there is any change of weight when iron rust is strongly heated in a crucible.
3. Find the action of hydrochloric acid upon iron rust? Is any Fixed Air evolved?
4. Repeat Experiment 2, using strong sulphuric acid instead of water



## CHAPTER XXII

### THE DISCOVERY OF OXYGEN

THE work of the Oxford chemists of the seventeenth century showed that common air was probably composed of two different gases, which we have called Fire-Air and Vitiated Air. No one however, had succeeded in isolating Fire-Air. It is evident, therefore, that though these chemists made a great advance towards a knowledge of the air and of combustion, more remained to be done. As a matter of fact, no further advance was made until the latter part of the eighteenth century through the investigations of *Scheele* in Sweden, of *Priestley* in England, and of *Lavoisier* in France.

### THE SEARCH FOR FIRE-AIR.

It must now be our object to isolate Fire-Air. We have found that when metals and phosphorus are heated in the air, they increase in weight, because they form compounds with the Fire-Air. Is it not possible that we may be able to separate the Fire-Air again from some one of these compounds? For instance, when metallic lead is heated for a long time in the air substances called *Red Lead* and *Litharge* may be obtained, which we are regarding as compounds of lead and Fire-Air. Now, is it not possible that by some treatment we may induce the lead and the Fire-Air to part company? An experiment in which this was effected was first made in 1774 by *Dr. Joseph Priestley*,<sup>1</sup> an eminent dissenting minister.

<sup>1</sup> Priestley says at the outset of his account of his experiments "more is owing to what we call *chance*, that is to the observation of events arising from unknown causes.

This is how he carried it out. He wrote :—

"Having procured a lens I proceeded with great alacrity to examine by help of it, what kind of air a great variety of substances would yield, putting them into the vessels represented (Fig. 93), which I filled with quicksilver and kept inverted in a basin of the same."



FIG. 93.

Amongst other substances he tried *red lead*. He found that an air was expelled from it, and at the same time the red lead became yellow like litharge. He got sufficient of the air to try a few experiments with it, and he says: "What surprised me more than I can well express was that a candle burned in this air with a remarkably vigorous flame,

and a piece of red-hot wood sparkled in it and consumed very fast."

We cannot readily repeat the experiment in just his way; for one reason, we cannot always have the sun's rays; but we can heat some red lead in a test-tube.

**EXPT. 1. A Fire-Air is obtained from Red Lead.**—Place some red lead in a narrow test-tube of hard glass. Heat it in the full flame of a foot-blowpipe. It at first darkens in colour and then fuses. Introduce into the tube a glowing splinter of wood, and notice that it bursts into flame.

It seems likely that the gas which the red lead evolves is the Fire-Air we are in search of. Note, too, that the residue is not a metal but is like the earth first obtained on roasting lead, called litharge. Try also whether a Fire-Air is given off when red lead is heated in a Bunsen flame, holding the tube as in Fig. 94.

#### PROBLEMS

Find whether any of the following substances yields Fire-Air when heated (1) in a Bunsen flame, (2) in a blowpipe flame:

Nitre	Pyrolusite. <sup>2</sup>
Chlorate of potash. <sup>1</sup>	Calx of tin.
Chalk.	Sand.

than to any proper design." He was ignorant of the work of the Oxford chemists, and after working on Fixed Air proceeded in a haphazard way to heat red lead and many other substances, simply to see what kind of air they would yield (*see Alembic Club Reprint, No. 7*).

<sup>1</sup> Fire-Air was obtained from chlorate of potash by Berthollet, about 1800.

<sup>2</sup> Also called manganese dioxide.

In the case of nitre compare what Scheele,<sup>1</sup> a poor Swedish apothecary, observed about the year 1772. He says: "I put an ounce of purified nitre in a glass retort for distillation and made use of a bladder, moistened and emptied of air in place of a receiver. As soon as the nitre began to glow it also began to boil and at the same time the bladder was expanded by the air that passed over. This is the cheapest and best method of obtaining Fire-Air."



FIG. 94.

In the course of his experiments Priestley heated a red substance called *calcined mercury*. It is formed upon the surface of the mercury when mercury is heated in common air. He found that an air was readily given off.

**EXPT. 2. Fire-Air is obtained from Calcined Mercury.**—Place a small quantity of calcined mercury<sup>2</sup> in a test-tube. Heat it cautiously but strongly in a Bunsen flame. It darkens in colour. Introduce a glowing splinter of wood. The wood bursts into flame. Observe also that a mirror of minute globules of mercury is deposited upon the cooler portion of the

<sup>1</sup> It should be noted that Scheele discovered Fire-Air before Priestley, about the year 1772. Priestley was ignorant of this (*see Scheele's Fire and Air*, Alembic Club Reprint, No. 8).

<sup>2</sup> Known to the chemist as oxide of mercury.

tube, which on being rubbed gather together into a drop of the liquid metal.

Thus calcined mercury when heated yields mercury and a gas which supports combustion very vigorously.

As we have said, Priestley discovered this Fire-Air in 1774 ; in the same year he told Lavoisier in Paris about it, and the famous French chemist at once repeated the experiment in a very instructive way.

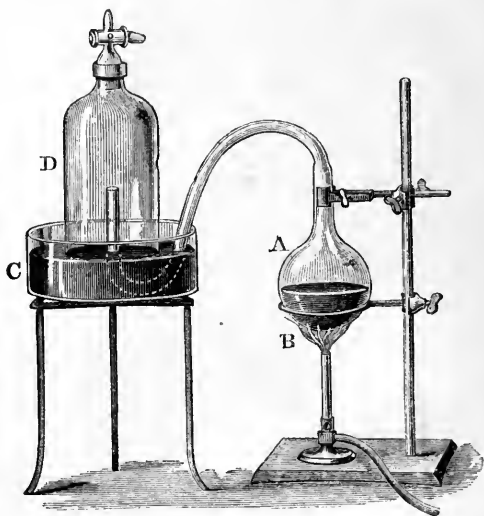


FIG. 95.

**Lavoisier's Experiments.**—(1) Lavoisier took a glass retort A with a long neck, and after bending the neck as represented in Fig. 95, he introduced some mercury, and placed the retort over a flame. The end of the neck came out above the surface of mercury in a dish C, and under a graduated bell-jar D containing 50 cubic inches of air. The mercury in the retort was heated strongly by the flame for several days. As this went on red specks of calcined mercury formed on the surface in the retort, while the mercury gradually rose in the jar. After a

few days, no further formation of the red substance appeared to be taking place, so after twelve days from the outset the experiment was stopped. It was found that the volume of the air in the jar had decreased between 7 and 8 cubic inches and that which remained in D was not fit for respiration nor did it support combustion. Then Lavoisier collected the red calcined mercury. It weighed 45 grains.

(2) The calcined mercury was next placed in a glass tube closed at one end, and bent as represented in Fig. 96, the open

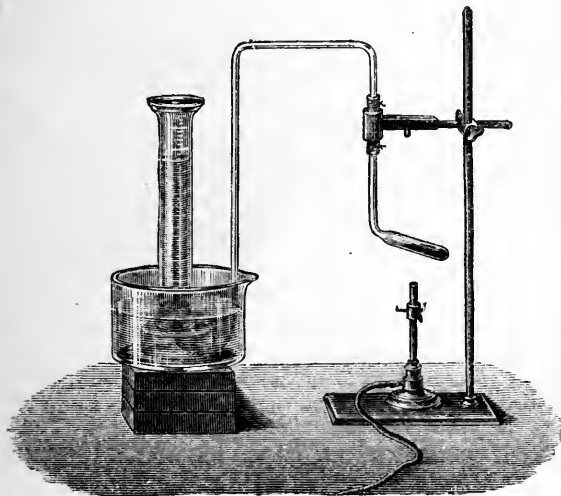


FIG. 96.

end being placed under a graduated cylinder filled with mercury. On heating the substance, a gas was given off, which measured between 7 and 8 cubic inches, and  $41\frac{1}{2}$  grains of mercury were deposited on the colder parts of the tube. The gas collected in the cylinder was proved to support vigorously the combustion of a taper or charcoal.

(3) Lavoisier repeated the analysis of 50 cubic inches of air as described above (1), and then added the 8 cubic inches of gas

obtained on heating the calcined mercury as described under (2) to the 42 cubic inches left in the bell-jar. He *obtained in this way an air which was in every respect like atmospheric air.*

Therefore, after having by *analysis* proved common air to contain in reality two airs or gases, he confirmed this by the method of *synthesis*,<sup>1</sup> that is, by putting the two airs together and so producing common air again.

It will be noticed that the volume of the air withdrawn by the mercury was about one sixth of the whole—not quite so much as we were able in a previous experiment to withdraw by means of phosphorus—because the action was not quite complete. Let it be particularly observed, however, that the volume of gas obtained from the calcined mercury by heat was the *same* as the volume of gas which disappeared from the bell-jar in Lavoisier's Experiment (1).

The experiments make it quite clear that the gas which calcined mercury evolves on being strongly heated is the very *same* which mercury picks out of the air, when calcined mercury is made. Therefore, the gas which calcined mercury gives off on being heated is the same as that portion of the air which we have called Fire-Air. We have thus found a method of preparing Fire-Air, and we can now proceed to try experiments in order to find out its characteristic properties.

Calcined mercury is an expensive substance, and we cannot therefore afford to use it to obtain sufficient Fire-Air in order to try many experiments. As soon, however, as Fire-Air had been discovered, other methods of preparing it were very soon found, and we will make use of one them. We have already found that chlorate of potash evolves Fire-Air when heated (Expt. 1, Problem). Try the effect of heating it when mixed with other substances.

**EXPT. 3. Preparation of Fire-Air from Chlorate of Potash.**—Heat the following substances in a test-tube in a Bunsen flame, using the apparatus shown in Fig. 97, and try to compare the rate at which Fire-Air is evolved.

- (1) Chlorate of potash.
- (2) Chlorate of potash mixed with a little pyrolusite.
- (3) Chlorate of potash mixed with a little sand.

<sup>1</sup> Greek : *syn*, together, and *thesis*, a placing.

**EXPT. 4. Laboratory Preparation of Fire-Air.**—Weigh out 25 grams of chlorate of potash and 16 grams of sand. Grind them together in a mortar. Introduce the mixture into a small flask provided with a stopper and delivery tube. Fill six jars

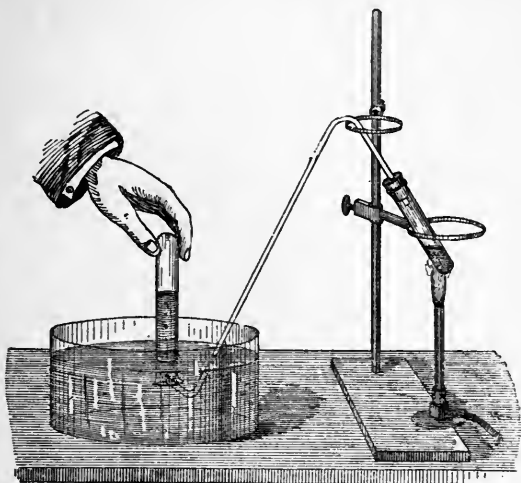


FIG. 97.

with water and invert them in a trough of water. Support the flask on wire gauze over a small flame and allow the first bubbles which rise through the water to escape. After collecting six jars of Fire-Air, remove the flask and delivery tube from the trough, and allow the flask to cool.

#### PROBLEM<sup>1</sup>

Find whether Fire-Air has the characteristic properties of Fixed Air.

**FIRE-AIR VIGOROUSLY SUPPORTS COMBUSTION.**—We may now proceed to test the characteristic properties of Fire-Air. Remember that it forms one fifth by

<sup>1</sup> NOTE.—The gas is now most conveniently obtained compressed in steel cylinders from Brin and Co.

volume of common air, and the chief difference between it and common air is that it is not mixed with Vitiated Air. What difference is this likely to make between the burning of phosphorus, sulphur, a candle or iron in it and in air?

**EXPT. 5. To burn a Candle in Fire-Air.**—Attach a small candle or a short taper to a deflagrating spoon or bent wire. Light it at a flame and introduce into a jar. It burns with a more brilliant white flame than in common air.

**EXPT. 6. To observe the Burning of Sulphur in Fire-Air.**—Place a piece of sulphur in a deflagrating spoon. Pour a little water into a gas jar. Hold the sulphur in a flame until it catches fire, and introduce it into the jar. The faint blue flame at once burns more brilliantly and white fumes are formed. When the flame dies out, remove the spoon, and observe the suffocating odour of the fumes. Cover the mouth of the jar with the hand and shake. The fumes are dissolved in the water. Add some litmus solution. It is turned red, showing that an acid has been formed. We may call it *sulphurous acid*.

**EXPT. 7. To observe the Burning of Phosphorus in Fire-Air.**—Place a little red phosphorus in a deflagrating spoon. Kindle it in a flame and introduce it into a jar. Observe that it burns with a very brilliant white light, and dense white fumes are produced, which readily dissolve in a little water. Add some litmus solution. It is turned red, showing the presence of an acid. We may call this *phosphoric acid*.

**EXPT. 8. To observe the Burning of Iron in Fire-Air.**—Fix a piece of a thin watch spring or fine clean iron wire to a deflagrating spoon. Dip the end of the spring in some burning sulphur. Introduce it into another jar in which some water (about an inch in depth) has been placed. The metal becomes red hot and then burns brilliantly, emitting sparks, and fused globules fall into the water. Examine these.

These experiments make it abundantly clear that those substances which burn in air burn much more brilliantly in Fire-Air. As Priestley expressed it, there is a "*superior goodness about this kind of air*." Some words Scheele wrote in 1777 also sum up the matter very well: "These experiments show, therefore, that this Fire-Air is just that air by means of which fire burns in common air; only it is there mixed with a kind of air



which places some hindrance in the way of the otherwise rapid and violent inflammation."

**Fire-Air is Respirable.**—We have previously come to the conclusion that Fire-Air is the only portion of the air which supports respiration and the life of animals. How then will pure Fire-Air suit them? Both Priestley and Scheele made experiments with animals and found that they were much more lively and lived more vigorously in Fire-Air than in common air. Here is an interesting passage Priestley wrote in 1775:—

"My reader will not wonder, that, after having ascertained the superior goodness of Fire-Air<sup>1</sup> by mice living in it, and the other tests above mentioned, I should have the curiosity to taste it myself. I have gratified that curiosity by breathing it. The feeling of it to my lungs was not sensibly different from that of common air; but I fancied that my breast felt peculiarly light and easy for some time afterwards. Who can tell but that, in time, this pure air may become a fashionable article in luxury. Hitherto only two mice and myself have had the privilege of breathing it!"

Lavoisier observed the same properties of the new gas, and he often called it "*vital air*" or "*eminently respirable air*."

It may be remarked that Priestley's prediction has in a measure been fulfilled, for at the present time Fire-Air is in use in our hospitals, when it is desired to help the breathing of patients.

**Lavoisier's Theory of Combustion.**—In the year 1774 when Scheele and Priestley had independently discovered Fire-Air, Lavoisier was aware that sulphur and phosphorus, or the metals tin and lead increased in weight when heated in the air. He put this down to common air getting "fixed" in them; but until he heard from Priestley of his discovery he did not thoroughly understand what it was that occurred in combustion. Then when he heard of the preparation of "Fire-Air," as Scheele called it, he made those experiments which we have described, and he saw clearly that in any case of combustion in the air whether of phosphorus or of a metal, or of a candle, or of anything else there was a combination taking place between it and that portion of the common air which we have called Fire-Air.

<sup>1</sup> Priestley gives it another name.

This was a wide statement to make, and one which applied to a great many cases. The Oxford chemists had 100 years before come to very much the same conclusion (*see* p. 166), but every one had forgotten their work. At the time of the discovery of Fire-Air almost every one believed in the Phlogistic Theory of Combustion and thought that when a substance was burnt a fiery principle *escaped* from it. In direct opposition to this Lavoisier maintained that in combustion the Fire-Air of the atmosphere *combined* with the substance burnt, so that the substance produced in the combustion was more complex, not simpler, than that which had been burnt. As a matter of fact Lavoisier's experiments and his clear statement of the case very soon convinced almost every one of the truth of his theory of combustion, and the former Phlogistic theory was abandoned.

**Combustibles and Supporters of Combustion.**—When a piece of phosphorus is burnt in Fire-Air, the phosphorus is said in common language to be the *combustible* substance, and the Fire-Air *the supporter of the combustion*. According to Lavoisier's theory, however, the combustion consists in the union of the two, and it is as correct to say that the phosphorus combines with the Fire-Air, as that the Fire-Air combines with the phosphorus. Take the case of coal-gas burning in air; it is usually said that the air is the supporter of the combustion, while the coal-gas is the combustible substance. Here we have two gases. If the combustion consists only in the union of the two, why should we call one the combustible substance any more than the other? The reason is that usually coal-gas is burnt *in* air. But *can* we burn air in coal-gas? Let us try.

EXPT. 9.<sup>1</sup> **To find whether Air will burn in Coal-Gas.**—Obtain a lamp glass. Fit the wider end with a cork in which are fixed a straight piece of wide glass tubing *a* (Fig. 98) about 10 cm. long, and also a narrower piece of glass tubing *c* bent at right-angles. Fit the upper end also with a cork in which is fixed a short glass tube *b*. Support the lamp glass in the ring of a retort-stand. Connecting the tube *c* with the coal-gas supply, pass coal-gas into the glass, and light the gas as it issues from *b*.

Then without turning off the gas, close the end of the tube *b* with a cloth, apply a flame to *a*, and at once remove the cloth.

<sup>1</sup> Lecture Table Experiment.

Notice that the flame which a moment before was burning at the end *a*, has travelled up the tube and is now burning at the other end. This flame, moreover, must be one of *air burning in coal-gas*.

It is seen therefore that the terms "combustible" and "supporter of combustion" are not used in a strictly accurate sense. They are however convenient.

The fact that air can be burnt in coal-gas just as coal-gas in air, supports Lavoisier's theory of combustion. Could it be explained by the former Phlogistic Theory?

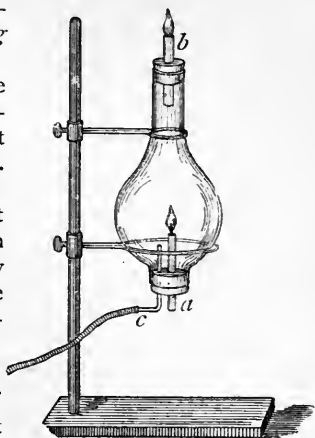


FIG. 98.

### THE COMPOSITION OF FIXED AIR.—

The student cannot fail to have observed that a piece of charred wood burns away and leaves hardly any ash. What has become of the charcoal? If we apply Lavoisier's theory of combustion, we must suppose that the charcoal or *carbon* as it is often called, has combined with Fire-Air, and formed some invisible air or gas. Since charcoal burns in air, it will probably burn all the more vigorously in pure Fire-Air. Let us try this.

**EXPT. 10. Fixed Air is produced where Carbon is burnt in Fire-Air.**—Take a large flask or cylinder and fill it with Fire-Air. Fix a piece of charcoal (charred bark best of all) in a deflagrating spoon by means of copper wire. Heat the charcoal in a flame until it is kindled, and then plunge it into the Fire-Air. It burns with great brilliancy.

With what different airs are we already acquainted? They are common air, Fire-Air, Vitiated Air, and Fixed Air. To see whether perchance any Fixed Air is formed, pour into the flask about 20 c.c. of clear lime water, cover the mouth with the palm of the hand and shake well. A turbidity is at once produced.

It seems then, that Fixed Air is produced when carbon is

burnt in Fire-Air. Fixed Air, therefore, is probably a compound of carbon and Fire-Air. If so, we should be able to separate carbon from Fixed Air. How is it likely we can do this? What substances can combine with Fire-Air, which we think there is in Fixed Air? Several metals do so, and magnesium especially readily. Let us try.

**EXPT. 11. Fixed Air contains Carbon.**—Fill a jar with Fixed Air. Fix a strip of magnesium ribbon to a deflagrating spoon. Kindle the end of the ribbon, and plunge it into the jar. The metal continues to burn. A white substance is formed like that produced when magnesium is burnt in air, but it is mixed with black specks, which look like charcoal or carbon. Add some dilute hydrochloric acid and shake. The calx of magnesium is dissolved, and the black specks of carbon are readily visible. Devise some means of proving that these specks are really carbon.

We have, therefore, confirmed by analysis our conclusion (arrived at by synthesis, Experiment 10) that *Fixed Air is a compound of carbon and Fire-Air*. Notice that it was the application of Lavoisier's theory of combustion which led up to this discovery.

**A new Name for Fixed Air.**—When we were first studying Fixed Air it was noticed that it possessed an acid character, for it was found to turn litmus solution red (p. 156). To express the fact that Fixed Air is an acid, and that it is a compound of carbon, Lavoisier proposed to call it *Carbonic Acid*—a name which has been in use ever since, and one which we shall now employ.

**A new Name for Fire-Air.**—We have found that when carbon is burnt in Fire-Air, carbonic acid is produced. The student will remember also that when sulphur and phosphorus are similarly burnt acids are also produced, which we have called sulphurous acid and phosphoric acid. These facts were observed by Lavoisier and led him to regard Fire-Air as a gas which had the property of producing acids when it combined with many substances. He therefore proposed to give Fire-Air the name *oxygen*—a word which means the *acid-producer*. It is derived from two Greek words *oxus*, acid, and *gennao*, I produce. This name gradually supplanted the old one, and is employed at the present day by every chemist for this gas.

At about the same time the name *nitrogen* was given to Vitiated Air, and it is the name now employed.<sup>1</sup>

**The Definition of an "Element."**—Many of the substances with which we have been dealing can no longer be regarded as simple in nature. The air around us contains two different gases, oxygen and nitrogen. Calcined mercury when heated liberates oxygen, and mercury is left. Red lead, again, is a compound of lead and oxygen. On the other hand, we have no evidence whatever that oxygen, or mercury, or lead are compound in nature, and we seem to require some word to distinguish such substances from those which are known to be more complex.

Such a word is at hand in the term *Element*, one which is in frequent use in ordinary language, and one which has for hundreds of years been employed by alchemists and chemists. But the Honourable Robert Boyle was the first to state clearly the sense in which it should be used. In his book *The Sceptical Chemist*, he pointed out in 1661 that only substances which can be obtained from compounds, and are not known themselves to be compounds should be regarded as elements.

**DEFINITION.**—*An Element is a substance which cannot by any known process be separated into two or more different substances.*

Then, lead, iron, and oxygen are, so far as our knowledge goes, examples of elements. Carbonic acid gas, on the other hand, is not an element, for we have been able to separate carbon from it, and we know moreover that it contains oxygen.

**The Definition of an Oxide.**—We have been led to recognise the presence of oxygen in the red calx of lead, in the white substance formed from burning phosphorus, and in the pungent smelling gas which arises from burning sulphur, as well as in carbonic acid gas. All these substances can be looked upon from the same point of view as compounds of oxygen with another substance, which, so far as our knowledge goes, is an elementary one. Thus, though they differ greatly from one another in appearance and in properties, yet from a chemical point of view they are all constituted in a similar way.

<sup>1</sup> This name was suggested by Chaptal in reference to the fact that Vitiated Air could be obtained from *nitre*.

It is convenient to call any such substance an *oxide*.

DEFINITION.—*An Oxide is a compound of an element with oxygen.*

A name which is often used for Fixed Air is *carbon dioxide*. The full significance of this cannot be understood at present, but it refers to the fact that the gas is a compound of carbon and oxygen.

#### TABLE OF SYNONYMS OF OXIDES.

It will be useful to tabulate for purposes of reference the old names and also the names which have been commonly adopted since the work of Lavoisier for some of the chief oxides :—

Litharge	Lead monoxide
Rust of iron	Ferric oxide
Calx of tin	Stannic oxide
Calx of zinc	Zinc oxide
Calcined mercury	Mercuric oxide
Calx of magnesium	Magnesium oxide, magnesia
Pyrolusite	Manganese dioxide
Fumes of burning phosphorus	Phosphorus pentoxide
Fixed Air	Carbon dioxide, carbonic acid
Fumes of burning sulphur	Sulphur dioxide

**Matter is indestructible.**—The thoughtful student has perhaps noticed that on more than one occasion in the preceding pages we have *assumed* that matter<sup>1</sup> is indestructible, and that it cannot be created. For instance, when we burnt charcoal in oxygen (Expt. 10, p. 191) we saw the charcoal disappear, but we took it for granted that it had not been annihilated. We assumed that it had combined with the oxygen forming some invisible gas, and we proceeded to try to find what this was and found it to be carbonic acid (Fixed Air). Then again, when we heated lead in the air (p. 163) and observed an increase in weight, we assumed that this could only be accounted for by some portion of the air becoming attached to the metal in other words, we did not regard the possibility of more lead having been created, but attributed the gain in weight to the addition of some external matter.

<sup>1</sup> DEFINITION.—*Matter* (or substance) is that which has (1) weight, (2) inertia (3) size.

Now the first impression of a child observing a candle disappear, or a small seed growing into a big plant, is that matter can be utterly destroyed, or, as in the case of the seed, grow and increase "of itself." Let us consider therefore what grounds we have for assuming, as we have done, that matter is everlasting and indestructible.

1. *Scheele's experiment* (p. 167) in which a piece of phosphorus was burnt in a closed flask full of air, verified this in a very clear way. For no change in weight had occurred after the combustion. This showed that although the phosphorus had burnt in part away, no loss of matter had occurred. Therefore, when the flask was opened under water, and it was found that one-fifth of the air had disappeared, we were perfectly justified in saying that this air had not been destroyed but had combined with the phosphorus forming a white solid substance. Thus, there were present in the flask :—

<i>Before the combustion</i>	<i>After the combustion</i>
Phosphorus	Unburnt phosphorus
Oxygen (Fire-Air)	Oxide of phosphorus
Nitrogen (Vitiating Air)	Nitrogen

In this case, therefore, the total weight of the substances after the chemical action was the same as the total weight at the outset.

2. Another illustration of the indestructibility of matter may be taken from *Lavoisier's experiment on the calcination of mercury* (p. 184). It was shown that when mercury was calcined in the presence of a given volume of air, the volume of the air was diminished by nearly one-fifth. This air was not destroyed. It had simply combined with the mercury forming a red substance. For when the calx of mercury was collected and heated it evolved exactly the same volume of gas again.

3. Again, a *burning candle* is apparently destroyed. But is this so? Let us test this.

EXPT. 12.<sup>1</sup> **To find whether there is a Loss of Matter when a Candle is Burnt.**—Take a round-bottomed flask of half a litre to one litre capacity. Select a tightly fitting rubber stopper *f*. Fix in this the rod *b* of a deflagrating spoon. Fix in it

<sup>1</sup> Lecture Table Experiment.

also a stout copper wire *a*, and a glass tube bent at right angles. These three must all be fitted *air-tight* in the rubber stopper.

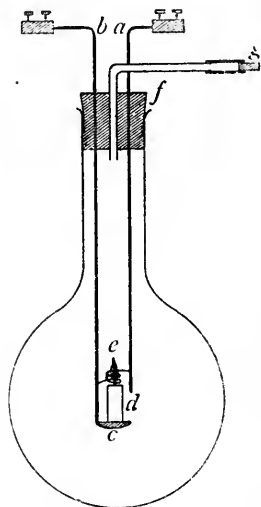


FIG. 99.

Fix a small piece of a candle in the cup *c*, and connect the rod bearing it with the copper wire by means of a platinum wire *e* wound twice round the wick. Fill the flask with oxygen, place the stopper tightly in the neck of the flask, and close the glass tube with a small piece of rubber tubing stopped with a short glass rod *g* (Fig. 99).

Support the whole at one end of the beam of a balance, and exactly counterpoise it. Then remove it from the balance to connect the rod and the copper wire for a moment with three Grove cells. The platinum wire *e* is heated white hot by the electric current, and the wick of the candle catches fire. The candle burns steadily for a time, and then the flame gradually dies down and goes out.

A portion of the candle has undoubtedly burnt away. Note the formation of dew on the sides of the flask. Has there been any loss of weight? Place the flask on the pan of the balance again with the same counterpoise. There is no change of weight. Then there has been no destruction of the matter of the candle.

These three instances will suffice to illustrate a law which is of the very highest importance in the study of chemistry, when the changes substances undergo are studied. This *law of the indestructibility of matter*, or *law of conservation of mass*, as it is often called, may be very briefly stated in the words:—

**MATTER CANNOT BE CREATED AND CANNOT BE DESTROYED.**

"We may lay it down," said Lavoisier, "as an incontestable axiom that in all the operations of art and nature, nothing is



created; an equal quantity of matter exists both before and after the experiment. Upon this principle the whole art of performing chemical experiments depends."

The discovery of carbonic acid (Fixed Air) did much to lead to a belief in the indestructibility of matter. It is evident that until it was found that carbonic acid is produced when charcoal is burnt, it was natural to suppose that the substance of the charcoal was actually destroyed. So, again, the law of the indestructibility of matter could not be grasped until the discovery of oxygen was made, and the part it played in combustion had been clearly set forth by Lavoisier.

Very truly did Scheele remark in 1777, "If the chemists of the preceding century had thought worthy of a more particular examination the elastic fluids resembling air which manifest themselves in so many operations, how advanced should we now be! They desired to see everything in corporeal form, and to collect everything as drops in the receiver. This is now for the first time better inquired into, and the air has begun to be carefully examined; and who is there that does not perceive the advantage which the results of such experiments carry with them?"

### EXERCISES AND PROBLEMS

#### The Discovery of Oxygen

1. Find whether any of the substances, A, B, C, D, E, F,<sup>1</sup> yield Fire-Air (Oxygen) when heated in a Bunsen flame.
2. Find which of the gas jars G, H, I, K, contain common air, Fire-Air, Vitiated Air, Fixed Air, respectively.

<sup>1</sup> e.g.  $\text{HgO}$ ,  $\text{BaO}_2$ ,  $\text{PbO}_2$ ,  $\text{K}_2\text{Cr}_2\text{O}_7$ ,  $\text{Fe}_2\text{O}_3$ ,  $\text{KClO}_3$ .

## CHAPTER XXIII

### THE ACTION OF ACIDS ON METALS.—THE DISCOVERY OF “INFLAMMABLE AIR”

EARLY in the seventeenth century it was reported that an inflammable air had been obtained by the action of sulphuric acid upon iron. The statement, however, attracted little or no attention, and until the time of the *Hon. Henry Cavendish*, in the next century, nothing further was known of any “Inflammable Air.” In the year 1766 this chemist published a paper entitled “Experiments upon Inflammable Air,”<sup>1</sup> in which he described the effect of treating different metals with the chief acids.

Let us set ourselves the same problem, and proceed to examine systematically the action of hydrochloric acid, sulphuric acid and nitric acid upon the following common metals:—zinc (granulated, or in strips of the sheet metal), iron (thin piano-forte wire or filings), tin (foil or granulated), lead (sheet or granulated), copper (turnings), and aluminium.

### THE DISCOVERY OF INFLAMMABLE AIR

#### ACTION OF HYDROCHLORIC ACID UPON METALS

**EXPT 1. To find the Effect of Hydrochloric Acid upon Zinc.**—Place two or three pieces of zinc in a test-tube. Add some hydrochloric acid. An immediate and rapid evolution of some air or gas takes place.

<sup>1</sup> Phil. Trans. of the Royal Society 1766, p. 144.

## PROBLEMS

1. Find whether strong hydrochloric acid diluted with three times its volume of water has any action on zinc.
2. Find the effect of hydrochloric acid (*a*) weak cold, (*b*) strong hot, on magnesium, tin, iron, copper, lead, aluminium. Tabulate your results.

Let us try whether the gas from zinc and hydrochloric acid is one of those we already know.

**EXPT. 2. The Gas from Zinc and Hydrochloric Acid is inflammable.**—Place some zinc in a test-tube. Add some weak hydrochloric acid. Introduce a lighted taper into the mouth of the test-tube. There is a slight explosion. Introduce the taper again. The gas burns with a blue flame. Can you suggest why it exploded at first?

The gas must be different from any gas we have had before, for the taper does not burn in it, as it does in atmospheric air or in oxygen; on the other hand the gas itself burns, being thus unlike nitrogen or carbonic acid. It was this property which led Cavendish to call it *Inflammable Air*. To study it further we must prepare it in larger quantity.

**EXPT. 3. Preparation of Inflammable Air.**—Place about 20 grams of zinc in a flask provided with a delivery tube and a thistle funnel (Fig. 100). Add water so as to cover the zinc and the end of the thistle tube. Add 25 c.c. of hydrochloric acid previously mixed with an equal volume of water. It has been shown that at first the gas is explosive. At this stage the gas in the delivery tube is like a train of gunpowder, the bottle like a bombshell. A light applied to the end of the tube would cause disastrous results. Collect the gas, a test-tube full at a time, over water. Try each test-tube full of it

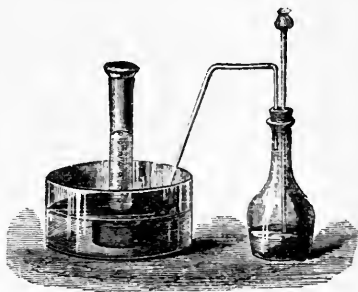


FIG. 100.

with a light. At first the gas explodes. The later portions catch fire gently, and the flame travels slowly down the tube. When it burns quietly collect the gas in larger quantities. Fill three jars. Half fill a test-tube with water; close it with the thumb, and invert it in the trough. Fill it up with the gas. It now contains air and the gas in equal parts. Set it aside.

The student will be better able to perform the above experiment with the least risk after studying some of the properties of the gas. (*See also* p. 201.)

### Properties of Inflammable Air.

**EXPT. 4. Inflammable Air cannot burn without Common Air.**—Light a taper. Hold a jar of the gas mouth downwards, and quickly plunge a taper into it. The gas burns quietly at the mouth of the jar, and the pale flame then travels upwards. The taper which was pushed into the centre of the jar is extinguished. As Cavendish remarked, "It appears that this air like other inflammable substances cannot burn without the assistance of common air."

This burning of Inflammable Air with a pale flame when only partially exposed to the atmosphere serves as a *test* to distinguish it from other gases.

Is it not likely, then, that the burning will be much more violent when the Inflammable Air *mixed* with common air?

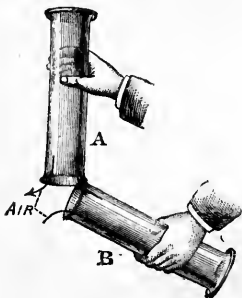


FIG. 101.

**EXPT. 5. Inflammable Air mixed with Common Air is Explosive.**—Apply a lighted taper to the test-tube set aside, containing the mixture. There is a sharp explosion, or as Cavendish expressed it, the mixture of the two airs "goes off with a bounce." We now can see

some reason for the precautions taken in Experiment 3.

**EXPT. 6. Inflammable Air is lighter than Common Air.**—(a) Place a jar containing the gas mouth upwards on the table. Remove the lid. After one minute insert a burning

taper. The gas does not catch fire. It must have escaped. So it is lighter than air.

(b) Hold a small strong jar *A* mouth downwards in the air. Try to pour the gas *upwards* into it from a jar *B* (Fig. 101). Apply a light to *A*. There is an explosion.

On account of the lightness of the gas it has been used for balloons. Soap bubbles blown with it will rapidly ascend through the air. Then will it not be possible to collect Inflammable Air by *upward displacement*? Try.

### PROBLEMS

Find whether it is the *same* gas—"Inflammable Air"—which is obtained by the action of

(a) Weak hydrochloric acid on iron, magnesium, aluminium.

(b) Strong hydrochloric acid on tin.

### ACTION OF WEAK SULPHURIC ACID UPON METALS.

**EXPT. 7. Weak Sulphuric Acid and Zinc evolve Inflammable Air.**—Place two or three pieces of zinc in a test-tube. Add some weak sulphuric acid. A gas is rapidly evolved. Apply a light. It burns. It is Inflammable Air.

### PROBLEM

Find the action, if any, of weak sulphuric acid upon aluminium, magnesium, iron, tin, copper, lead.

Inflammable Air is commonly prepared by the action of weak sulphuric acid upon zinc in a similar way to that described in Experiment 3.

### Precautions in the Preparation of Inflammable Air.

—(1) The gas is often made in a Woulff's bottle of thick glass. Though this is not readily broken by heavy pieces of zinc falling into it, thick glass is specially apt to crack on heating. It is important, therefore, to keep the contents *cool*.

(2) See p. 135, Experiment 3, on mixing sulphuric acid and water. After mixing 8 parts of water with 1 part of sulphuric acid, set the mixture aside to cool; otherwise it might crack the Woulff's bottle. It is better still to keep a stock of diluted acid ready.

(3) Make sure that all air is expelled before the gas is collected (*see* p. 199, Experiment 3).

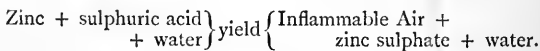
(4) Should the contents of the bottle become hot, the current of gas will probably slacken. It is then best to pour off the liquid, wash the zinc, and start afresh.

(5) Take care not to breathe the gas made from ordinary impure zinc.

(6) The addition of a few drops of a solution of copper sulphate often quickens the current of gas.<sup>1</sup>

**EXPT. 8. To find what becomes of the Zinc in the Preparation of Inflammable Air.**—Zinc placed in excess of dilute sulphuric acid entirely disappears. What becomes of it? Fill a beaker one-third full with dilute sulphuric acid. Put in several pieces of zinc. A vigorous effervescence goes on and gradually subsides. Why does the effervescence cease? Is the sulphuric acid all spent? Or the zinc all used up? Add more zinc. Effervescence begins again. This shows there is still some acid in the liquid. Go on adding zinc from time to time until there is no more action, and the sulphuric acid has spent itself. The liquid now contains undissolved zinc, and floating black specks, which are impurities left by the dissolved zinc. Filter. Evaporate and crystallise.

The long white crystals obtained are called *zinc sulphate*, because they are obtained from zinc and sulphuric acid. Thus—



#### PROBLEM

Try to prepare crystals of (1) iron sulphate, (2) magnesium sulphate, by dissolving iron and magnesium respectively in sulphuric acid.

**What is the Source of Inflammable Air?**—Does the Inflammable Air come from the metal? We have obtained it by using several metals; then, if it comes from the metals, they are not elements. Or does it come from the acid? If so, do other acids contain it as a common constituent? Or does

<sup>1</sup> See Appendix for a description of Kipp's Hydrogen Apparatus.

the Inflammable Air come from both the metal and the acid? Does it perhaps come from the water? And is it an element or a compound?

We cannot as yet answer these questions. We must wait for further evidence, and must in the meantime endeavour to avoid stating that the gas is obtained from the metal.

**Measurement of the Volume and Weight of Inflammable Air obtainable by the Action of Acids on Metals.**

**EXPT. 9. To find the Volume of Inflammable Air obtainable by means of One Gram of Magnesium.** — Fit up the apparatus shown in Fig. 102, consisting of funnel, rubber tube, spring clip, burette, and wide glass dish or beaker. The rubber tube *b*

must fit the burette tightly. Take 5 cm. of magnesium ribbon. Its weight will be only a few centigrams, and a single milligram will make a large percentage difference in the result. Weigh, if possible, exactly to a milligram. Disconnect the rubber tube from the burette. Gently bend the magnesium. Push it one-third of the way up the burette to *a*. Suck up water in the burette to fill it. Close the tap. Re-attach the rubber tube. Support the whole by the ring of a retort-stand. Half fill the funnel with dilute hydrochloric acid. Open the clip. Gently open the tap. Acid runs through. As soon as bubbles start from the magnesium turn off the tap. The magnesium is quickly and almost entirely dissolved. The burette must now be slightly tilted and a few more drops of acid skilfully run down the side to dissolve the rest of the metal. Read the position of the liquid in the burette.

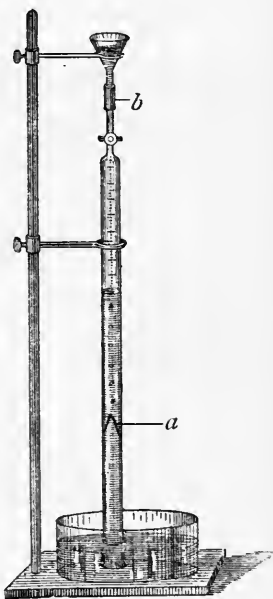


FIG 102

This is not the *volume* of the gas, because the graduations do not count from the tap. Devise some means of finding this volume, say 23 c.c.

Thus, in an experiment,

0.25 gram magnesium gave with acid 23 c.c. Inflammable Air.

Hence, 1 " " would give with acid 920 c.c. " "

Do you find that magnesium gives with acid the same volume of gas?

**EXPT. 10. To find the Volume of Inflammable Air obtainable by means of One Gram of Zinc.**—Fit up the apparatus shown in Fig. 103, consisting of a 4 oz. flask, Winchester

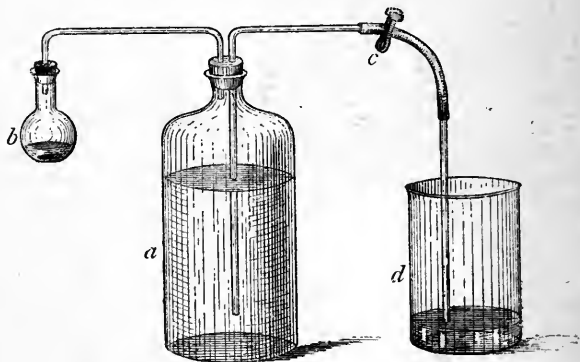


FIG. 103.

pint bottle, beaker, rubber corks, glass tubes, rubber junctions, and clip. The rubber corks must fit the flask and the bottle tightly. Nearly fill the bottle with water. Open the clip *c*. Suck at the end of the tube *d* so as to fill it with water, and then close the clip *c*. The tube will remain full of water. Place an empty beaker under *d*. Fill the flask *b* half full of dilute sulphuric acid (1 of acid to 3 of water).

Weigh out accurately about 1 gram of zinc. Wrap it up in a small piece of paper. Remove the flask *b* attached to the cork, introduce the zinc, and *at once* replace the cork, and open the clip *c*. The gas evolved as the metal is dissolved enters the Winchester bottle, and forces an equal volume of water through:



the tube into the beaker. If the solution of the metal slackens or stops, the flask *b* must be heated over wire gauze.

When the metal is completely dissolved the flask *b* and its contents are found to be hot, and must be cooled. Why? This can be readily done by placing the flask in a beaker of cold water. Next, raise the beaker until the level of the water in it is the same as that of the water in the bottle. Then close the clip *c*. Now measure the volume of the water in the beaker by pouring it into a measuring jar. This is also the volume of gas evolved on the solution of the metal.

Thus, 951 gram zinc gave with acid 350 c.c. Inflammable Air.

Hence, 1    ,,       ,,       ,,       368 c.c.       ,,

### PROBLEMS

(1) Find the volume of Inflammable Air yielded by one gram of zinc with dilute hydrochloric acid (1 of acid to 2 of water).

(2) Find what volume of gas can be obtained from one gram of iron with (a) dilute sulphuric acid, (b) dilute hydrochloric acid.

(3) Find the volume of gas obtainable from one gram of magnesium with (a) dilute sulphuric acid, (b) dilute hydrochloric acid.

(4) Find the volume of gas obtainable from one gram of tin with strong hydrochloric acid.

Cavendish found that (a) the *same* volume of Inflammable Air was obtained whether sulphuric or hydrochloric acid was used, (b) the same weight of different metals gave *different* volumes of Inflammable Air. Have you arrived at the same result? He found—<sup>1</sup>

1 gram tin yielded with acid 202 c.c. Inflammable Air.

,,    zinc    ,,       ,,       356 c.c.       ,,

,,    iron    ,,       ,,       412 c.c.       ,,

**EXPT. II. To find the Weight of Zinc by means of which One Gram of Inflammable Air can be obtained.**

—Fit up the apparatus shown in Fig. 104, consisting of a 4 oz. flask, "drying tube," cork, and bent glass tube. The flask must be light, and of thin glass. Pack the tube with dry calcium

<sup>1</sup> His results, given by him in English measure, have been stated above in the units of the metric system.

chloride. Fill one quarter of the flask with dilute sulphuric acid. Add one or two drops of a solution of copper sulphate. Weigh the whole on a good balance. Weigh out accurately 0.5 to 1 gram of pure zinc foil.

Remove the cork, introduce the zinc, and at once replace the

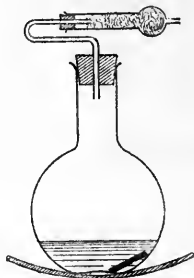


FIG. 104.

cork. If the evolution of gas becomes rapid, the flask must be cooled by placing it in cold water. When the zinc is completely dissolved, dry the outside of the flask with a cloth and weigh again. Thus :—

Weight of flask and acid	... ..	= 98.689 grams.
Weight of zinc	... ..	= 0.971 „
Then, Weight of flask, acid, and zinc	... ..	= 99.660 „
Final weight of flask, &c.	... ..	= 99.628 „
Hence, Weight of Inflammable Air evolved...	... ..	= 0.032 „

Thus,

0.032 gram Inflammable Air was obtained by means of 0.971 gram zinc.

Hence,

1 gram Inflammable Air was obtained by means of 30.3 grams zinc.

It will probably be difficult to obtain an accurate result on account of the very small weight of the Inflammable Air. Still, it was in almost exactly this way that Cavendish tried to answer the same question.

#### EXERCISE

Work out in a similar way the results of your previous experiments, assuming 1 litre of Inflammable Air to weigh 0.09 gram,

## CHAPTER XXIV

THE ACTION OF ACIDS ON METALS.—THE DISCOVERY OF  
VOLATILE SULPHUROUS ACID AND OF NITROUS AIR

### THE DISCOVERY OF VOLATILE SULPHUROUS ACID

ACTION OF STRONG SULPHURIC ACID UPON METALS.—  
This is the next question to be considered in continuing our  
research upon the action of acids upon metals.

EXPT. 1. **To find the Effect of Strong Sulphuric Acid  
upon Zinc.**—Place a piece of zinc in a test-tube. Cover it  
with strong sulphuric acid. A few bubbles are given off.  
Warm. The bubbling increases. A gas is given off. Test it  
with a light. A taper is extinguished in it, and the gas does  
not catch fire. It is not Inflammable Air. It has a strange  
pungent sulphurous smell. A new gas is being formed.

#### PROBLEM

Find the effect of hot strong sulphuric acid upon copper, iron, tin,  
mercury, lead.

Cavendish observed in 1766 that on heating zinc, iron, or tin  
with strong sulphuric acid, there was given off what he called  
*Volatile Sulphurous Acid*.<sup>1</sup> He did not collect it, nor examine

<sup>1</sup> It is now known as sulphur dioxide.

it. This was first done by Priestley in 1775. He collected the gas over mercury (we shall see presently why he did not collect it over water). Since the gas was obtained by the action of strong "vitriolic acid" on metals Priestley called it *vitriolic acid air*.

It is best prepared by means of copper and strong sulphuric acid.

**EXPT. 2. Preparation of Volatile Sulphurous Acid.**—Introduce 10 grams of copper into a flask. Fit it with a stopper

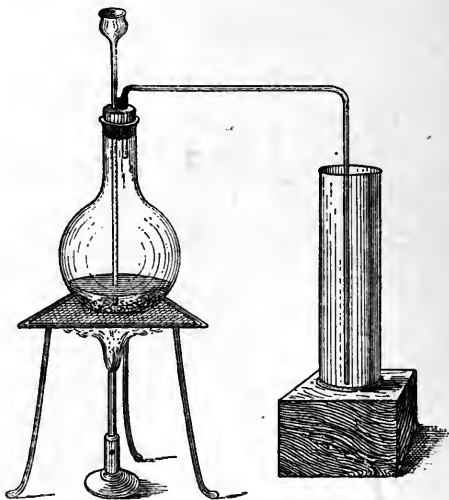


FIG. 105.

through which are passed a thistle funnel and a delivery tube bent twice at right angles (Fig. 105). Support the flask upon wire gauze on a tripod stand, and place a *dry* jar so as to collect the gas by downward displacement. Add 50 c.c. of strong sulphuric acid. Heat cautiously. Plunge a burning taper from time to time into the jar. As soon as it is extinguished at the mouth of the jar, remove the jar, closing it with a stopper or glass plate, and substitute another *dry* jar.

## Properties of Volatile Sulphurous Acid.

We have observed that the gas is not inflammable, and that it does not support combustion. It is heavier than air (for we were able to collect it by downward displacement). It possesses a suffocating sulphurous odour, unlike any other gas we know. Why could we not collect it over water?

**EXPT. 3. It is soluble in Water.**—Invert a jar filled with the gas in water, and withdraw the glass plate closing its mouth. Shake the jar so as to wet the sides. The water *ascends* in the jar. Why? The gas must be soluble in water. Then the water has perhaps become acid or alkaline? Test with litmus. The solution of the gas in water is acid.

### PROBLEM

If there be any residue left in the jar after shaking, find out what it is.

**EXPT. 4. Properties of the Aqueous Solution of the Gas.**—Half fill a test-tube with water. Pass a current of gas to the bottom of it. Notice that the bubbles of gas do not reach the surface. (Water, at the ordinary temperature, can absorb about 40 times its own volume of the gas.) Divide the liquid into three parts:—

(a) Smell it; it has the same odour as the gas. Taste it: it is acid.

(b) Add blue litmus; it is reddened.

(c) Add lime water; there is no turbidity.

**EXPT. 5. It bleaches many Colouring Matters.**—(a) Pour a dilute solution of the dye magenta into a jar containing Volatile Sulphurous Acid. Cover the mouth with the hand and shake well. The colour is nearly or entirely destroyed.

(b) Place some red or blue flowers in a jar of the gas. The colours are soon destroyed.

The gas is used for bleaching straw, silk and woollen goods.

We must for the present leave this bleaching action unexplained.

Volatile Sulphurous Acid is found also to be a disinfectant. Rooms, bedding and clothing which contain disease germs are therefore often exposed to the action of the gas.

## PROBLEM

Find whether the gas given off by the action of strong sulphuric acid upon (a) mercury, (b) zinc, is Volatile Sulphurous Acid.

**The Composition of Volatile Sulphurous Acid.**—Priestley quite expected that when Volatile Sulphurous Acid was dissolved in water sulphuric acid would again be obtained. But on the contrary, as he himself wrote in 1775 :—

“Water impregnated with Volatile Sulphurous Acid differ remarkably from sulphuric acid.<sup>1</sup> Its acidity is now become trifling to what it was, and from being the most fixed and strongest, it is now become the weakest and the most volatile of all acids—the smell of it being intolerably pungent, and almost the whole of it evaporating when it is exposed to the open air.”

It is therefore evident that the gas does not occur in solution in ordinary sulphuric acid.

The odour of Volatile Sulphurous Acid, however, is very similar to that produced when sulphur is burnt in oxygen or in air. Perhaps the *same* gas is produced by burning sulphur as by the action of strong sulphuric acid upon metals?

**EXPT. 6. To find whether the burning of Sulphur in Oxygen also produces Volatile Sulphurous Acid.**—Place some sulphur in a deflagrating spoon. Light it and plunge it into a jar of oxygen. Compare the properties of the contents of the jar with those of Volatile Sulphurous Acid. They are the same.

## PROBLEM

Devise means for heating some sulphur in a glass tube in a stream of air, and passing the products of the burning sulphur into water.

Then the Volatile Sulphurous Acid evolved by the action of strong sulphuric acid upon metals is really an *oxide of sulphur*. It is therefore a compound, and not an elementary substance. It is commonly called *sulphur dioxide*.

Have we any evidence as to the *source* of the gas?

Let us call the acid which is formed when the gas is dissolved in water *sulphurous* acid and so distinguish it from *sulphuric* acid.

<sup>1</sup> His own words here are “vitriolic acid air” and “vitriolic acid.”

## THE DISCOVERY OF NITROUS AIR

### ACTION OF NITRIC ACID UPON METALS

EXPT. 7. To find the Effect of Cold Nitric Acid upon Zinc and Copper.—(a) Place about 1 gram of zinc in a small flask. Mix 5 c.c. strong nitric acid with 5 c.c. water. Cool the mixture. Pour it into the flask. Is there any action? Very little.

(b) Repeat with copper. There is an effervescence. Copious red fumes are evolved. Test it with a light. It is extinguished, and the gas does not burn. There is a characteristic smell. It forms red fumes in the air. Yet another new gas is being formed.

### PROBLEM

Find the effect of diluted nitric acid upon iron, tin, lead.

You have been repeating experiments which were made more than 100 years ago. Cavendish, in 1766, tried the action of nitric acid upon zinc, iron, and tin. When giving an account of his experiments, all he wrote was:—

“All these metallic substances dissolve readily in the nitric acid and generate air, but the air is not at all inflammable.”

Priestley was the first to collect the gas and examine its properties. This is what he wrote:—<sup>1</sup>

“On the 4th of June, 1772, I first found this remarkable species of air. . . . When I first began to speak and write of it to my friends, I happened to distinguish it by the name of ‘Nitrous Air’<sup>2</sup> because I had procured it by means of spirit of nitre.”

Priestley wrote that he also obtained Nitrous Air by the action of nitric acid upon iron, tin, and lead, but none in the case of zinc. Further, he adds:—

“I have generally found it most convenient to get Nitrous Air from copper, on account of the pretty equable solution of the metal in the nitrous acid. If iron is made use of, the process is much more difficult.”

<sup>1</sup> *Experiments and Observations on Different Kinds of Air*, Dr. J. Priestley 1790 edition, vol. 1, p. 328.

<sup>2</sup> It is now known as Nitric Oxide.

The method of preparation by means of copper is still commonly used.

**EXPT. 8. Preparation of Nitrous Air.**—Fit up the apparatus in Fig. 100, with a glass delivery tube. Place 30 grams of copper turnings in the flask. Mix 50 c.c. strong nitric acid with 50 c.c. water. Cool. Pour the acid down the funnel so as to cover the end of the tube. Red fumes are produced. At first little gas comes over. Then a colourless gas is collected. Neglect the first two jars. Collect three more. If the action slackens, add 1 or 2 c.c. strong nitric acid.

#### PROBLEM

What became of the copper in the last experiment? Try to obtain crystals of copper nitrate.

### Properties of Nitrous Air.

Nitrous Air is a colourless gas, insoluble in water, and we found it did not support the combustion of a taper. Let us try this again.

**EXPT. 9. Nitrous Air does not support Ordinary Combustion.**—Plunge a lighted taper into a jar of the gas. It is extinguished.

#### PROBLEM

Find whether sulphur and phosphorus kindled in a deflagrating spoon will burn in nitrous air.

Observe that red fumes always appear about the mouth of the jar when it is exposed to the air, having a characteristic disagreeable smell.

What is the cause of these red fumes? Is it not probable that the red fumes are produced by the Nitrous Air mixing with common air? Let us test this.

**EXPT. 10. Nitrous Air forms Red Fumes with Common Air.**—Fill a jar one-third full with water; invert it over water so that it is two-thirds full of air. Bubble up into it one-third of a jar of the Nitrous Air. Red fumes at once appear. Shake. The water gradually rises, and at the same time the



red fumes disappear. Finally the two parts of common air and one part of Nitrous Air yield rather less than two parts of gas.

What has become of the rest? It must have dissolved in the water. Push a strip of blue litmus paper up into the water in the jar. It is reddened. Then the red fumes have dissolved in the water, forming an *acid* solution. Here is what Priestley wrote of the experiment:—

“I hardly know any experiment that is more adapted to amaze and surprise than this, which exhibits a quantity of air which as it were devours a quantity of another kind of air half as large as itself, and yet is so far from gaining any addition to its bulk that it is considerably diminished by it.”

Have we completely determined the cause of the red fumes? No. Common air contains both oxygen and nitrogen, so we must find whether the red fumes are due to the action of the nitrogen or of the oxygen, or of both, upon the Nitrous Air.

**EXPT. 11. Does Nitrous Air form the Red Fumes with Nitrogen?**—Obtain some nitrogen as in Experiment 12, p. 170. Open the flask under water and shake. Bubble the nitrogen up into a jar (standing over water) containing a little Nitrous Air. *No* red fumes result, and the volume of gas is increased.

**EXPT. 12. Does Nitrous Air form Red Fumes with Oxygen?**—Fill two-thirds of a jar with oxygen. Add to it one-third of a jar of Nitrous Air, as in Experiment 10. Priestley first tried this experiment in 1775. He found that the volume was diminished and “the redness was really deeper and the diminution something greater than common air would have admitted.”

Add another third of a jar of Nitrous Air. There is a further diminution in volume.

Then, the red fumes produced when Nitrous Air is added to common air, are due to the Nitrous Air forming a red coloured gas with the *oxygen* of the air, which red gas is readily dissolved in water, forming an acid solution.

#### EXERCISE

Find whether Nitrous Air forms red fumes with (a) Inflammable Air, (b) Fixed Air.

**Nitrous Air may be used as a Test for Oxygen.**—Since Nitrous Air forms red fumes with oxygen and not with any other gas, we are provided with a *test* to see whether a gas is oxygen, or has oxygen mixed with it.

Further, from the amount of contraction when the gases are shaken with water, we can form some idea of the amount of oxygen present. This test was formerly largely used by chemists after its discovery by Priestley. Cavendish showed, however, that it is not possible to obtain accurate results with simple apparatus, nor without observing many points of detail.

John Dalton was accustomed to analyse air in this way. On giving, in 1802, an account of his experiments, he wrote :—<sup>1</sup>

“ In all parts of the earth and at every season of the year the bulk of any given quantity of atmospheric air appears to be reduced nearly 21 per cent., by abstracting its oxygen. Air brought from the summit of Helvellyn in Cumberland (1,100 yards above the sea) gave no perceptible difference from the air in Manchester.”

**What is the Composition of Nitrous Air?**—Is it a compound or an element? We found that vigorously burning phosphorus would continue to burn in Nitrous Air (Experiment 9, Problem). This suggests that Nitrous Air may contain oxygen. How can we make sure? Refer back to p. 191, Experiment 10. Let us see whether a metal, magnesium, for example, can extract oxygen from Nitrous Air.

**EXPT. 13. To find whether Nitrous Air contains Oxygen.**—Take a Ramsay burner (Fig. 106).<sup>2</sup> This will give a long horizontal flame over which a glass tube can be supported. Take a hard glass tube *ab*, about 12 inches in length. Place in it a few strips of magnesium ribbon, and support it over the Ramsay burner. Pass Nitrous Air through the tube, drying the gas first by bubbling it through strong sulphuric acid in a flask *c*. After sweeping out the air with the gas, light the burner. Presently the magnesium catches fire, and a white calx is left. What is this? It is magnesium oxide. Then Nitrous Air must contain oxygen.

<sup>1</sup> See Alembic Reprint, No. 2, p. 8.

<sup>2</sup> Called after Professor Ramsay, now of University College, London.

## PROBLEM

Find the effect of passing Nitrous Air over heated (*a*) copper, (*b*) iron.

But Nitrous Air must contain something else besides oxygen. Suppose we remove the oxygen as in the last experiment, and endeavour to collect any gas that may issue from the tube over water.

EXPT. 14. To find what Nitrous Air contains besides Oxygen.—Pack the tube *ab* with clean copper turnings. Attach a delivery tube to the end *b*. Expel air from the apparatus by means of the gas, and then light the burner.

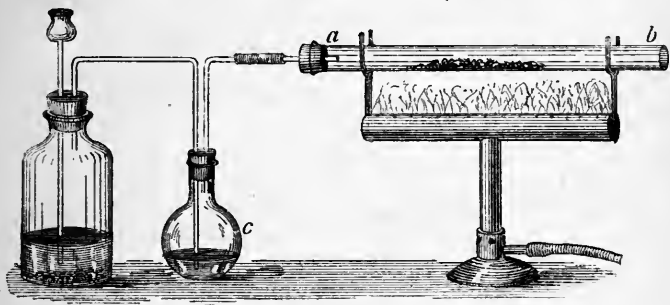


FIG. 106.

Collect the gas that issues in jars over water. Neglect the first jar. Examine the gas collected in the second jar.

The gas no longer reddens on contact with the air. Apply a light. It goes out. Is it carbonic acid gas? Add lime-water. It remains clear. Then the gas is nitrogen.

Nitrous Air, therefore, contains both nitrogen and oxygen. But so does atmospheric air. Is it the same as air? No, for it extinguishes a taper, and it forms red fumes with air or with oxygen. Then, in what does the difference between air and Nitrous Air consist? This question must be left unanswered at present.

Nitrous Air is now called *Nitric Oxide*, a name which denotes the fact that it is a compound of nitrogen and oxygen. Observe that the name would be equally applicable to atmospheric air, so far as we know at present.

### THE ACTION OF AQUA REGIA UPON METALS

Every metal we have tried has dissolved in at least one or other of the acids, nitric, sulphuric, or hydrochloric. Can gold be dissolved by these acids?

EXPT. 15. **Gold is soluble in Aqua Regia only.**—Half fill a test-tube with strong hydrochloric acid. Into a second test-tube pour one-third of this volume of strong nitric acid.

(1) Breathe on a glass rod. Take up half a sheet of gold leaf and put it in the nitric acid. Warm gently. Stir. It breaks up into little flakes but does not dissolve.

(2) Repeat with the hydrochloric acid. The gold is again not attacked.

(3) Mix the liquids. The gold dissolves readily.

This mixture of hydrochloric and nitric acids is called *Aqua Regia*, because it dissolves the royal metal gold.

### PROBLEMS

1. Find what acids will dissolve platinum.
2. Find whether the given  $\left\{ \begin{array}{l} \text{coin} \\ \text{metal} \\ \text{gilt leaf} \end{array} \right\}$  is gold or tinsel.

### EXERCISES

#### Preparation of Gases

Try the following experiments at first in test-tubes. Examine the properties of the gases formed. You are not expected to identify them in every case. Devise suitable apparatus for repeating the experiments on a larger scale.

1. Heat a mixture of bleaching powder, water, and copper oxide.
2. Heat strong sulphuric acid and powdered charcoal.
3. Heat strong nitric acid and a strong solution of ferrous sulphate.
4. Heat a strong solution of caustic soda with some zinc and a little copper sulphate.
5. Heat zinc and dilute nitric acid.
6. Heat ammonium chloride and caustic soda.
7. Heat ammonium chloride and sulphuric acid.
8. Heat ammonium nitrate.
9. Heat lead nitrate.
10. Heat oxalic acid and strong sulphuric acid.

## CHAPTER XXV

### THE DISCOVERY OF THE COMPOSITION OF WATER

#### (A) THE SYNTHESIS OF WATER

WHEN Inflammable Air is mixed with common air and fired, it explodes, or as Cavendish put it, it goes off "with a bounce." What will happen if it is mixed with oxygen instead of common air?

EXPT. 1. To show that Inflammable Air explodes much more violently with Oxygen than with Common Air.—Fill a sound soda-water bottle with water; invert it in water. Bubble Inflammable Air up into it, until two-thirds of the water is displaced. Fill it up with oxygen. Close the neck with a cork. Wrap a cloth round the bottle. Remove the cork and quickly hold the neck to a flame.

This experiment was first tried by Priestley about 1774, and he remarked, "The effect far exceeded my expectations, and it has never failed to surprise every person before whom I have made the experiment. The report will be almost as loud as that of a small pistol; being, to judge by the ear, not less than 40 or 50 times as loud as with common air."

Can you suggest any reason for the explosion being so much more violent in the one case than the other? The explosion is so rapid that we can hardly say what has occurred. Let us try to burn Inflammable Air *slowly* in a closed vessel containing common air or oxygen.

EXPT. 2. To show that a Contraction of Volume occurs when Inflammable Air is burnt in Common

**Air or in Oxygen.**—Select a wide-necked flask *a* of about 500 c.c. capacity (Fig. 107). Fill a large deep beaker or trough with warm water. Place some zinc in a small (2 oz.) flask *b* and add dilute sulphuric acid, so that Inflammable Air is rapidly evolved. Close it with a cork provided with a short upright delivery tube *c*, the end of which has been narrowed by holding it in a Bunsen flame. Light the Inflammable Air at the opening of the tube, observing the necessary precautions (*see* p. 201). Hold

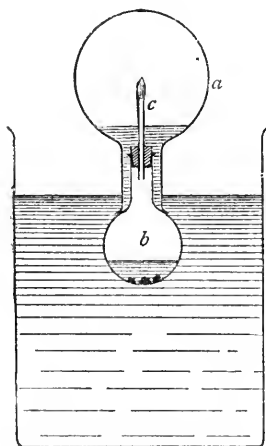


FIG. 107.

the small flask *b* in the water, with the flame just above the surface, and invert the larger flask *a* over the flame, so that its mouth is under the water. Then, releasing the flask *b*, the upward pressure of the water will cause it to take the position shown.

The Inflammable Air continues to burn, but the flame soon becomes fainter and smaller. As soon as it is seen to be extinguished quickly remove the flask *b* by lowering it in the water. Observe *that the water has risen in the flask a*. Then, some of the air has disappeared. Gently shake the flask *a* to throw the water against the top in order that the flask and the residual air may acquire the tem-

perature of the water. Close the mouth of the flask *a* by a cork under water. Remove the flask *a* and pour its contents into a measuring cylinder. Find also the total capacity of the flask *a*.

To quote an actual experiment :—

Volume of water which entered the flask	...	105 c.c.
Total volume of the flask	...	490 c.c.

Hence,

Contraction in the volume of the air in the flask = 21.4 per cent

You also have probably found a contraction of the air. How much?

## PROBLEM

Repeat the experiment with the difference that a longer delivery tube is used, and the flask is filled with oxygen in place of air. Find what volume of the oxygen is lost.

These experiments were made by Scheele (about 1773). He found that on using common air, the *fifth* part of it was lost. In the case of oxygen, nearly the *whole* was lost.

Can you draw any conclusions from the results of these experiments? In each case the oxygen present seems to disappear. What has become of it? Unless our theory of the conservation of mass is at fault, the oxygen cannot have been destroyed.

**EXPT. 3. To show that a Dew is formed when Inflammable Air is burnt.**—Take a clean dry test-tube. Introduce a little Inflammable Air into it by upward displacement. Hold its mouth to a flame. It explodes. Carefully examine the sides of the tube.

Priestley did this, and found that “the inside of the glass, though clean and dry before, immediately became *dewy*.”

Where has the dew come from? Is it merely the moisture which we know is present in common air?<sup>1</sup> If so, we still have not accounted for the Inflammable Air and oxygen which disappeared in Experiment 2. It is evidently important to see whether there is after all any *loss of weight*. This task the Hon. Henry Cavendish set himself, when he heard of Priestley’s experiment.

**Cavendish’s Experiment to show that there is no Loss of Weight on firing a Mixture of Common and Inflammable Air.**<sup>2</sup>—We cannot readily repeat Cavendish’s experiment. He used a strong glass globe. He filled it several times with different proportions of common and Inflammable Air. The mixture was fired by electricity. He found *no loss of weight*. In all the experiments the inside of the glass globe became dewy. He also succeeded in measuring the diminution in volume produced by the explosion.

<sup>1</sup> This is the conclusion Priestley and Warrtore came to in 1781. See *Experiments on Air*, by the Hon. Henry Cavendish, Alembic Club Reprint, No. 3, p. 12.

<sup>2</sup> *Loc. cit.*

Cavendish concluded that "when inflammable and common air are exploded in a proper proportion, almost all the inflammable air and near *one-fifth* (*see* Experiment 2) of the common air, lose their elasticity and are condensed into dew."

What is the nature of this dew? We must obtain a large amount of it.

EXPT. 4.<sup>1</sup> **To collect the Dew formed on burning a Mixture of Common and Inflammable Air.**—Fit the tubulure of a retort *a* with a cork, through which is passed a piece of glass tubing *b*, as shown in Fig. 108. Fix the retort in

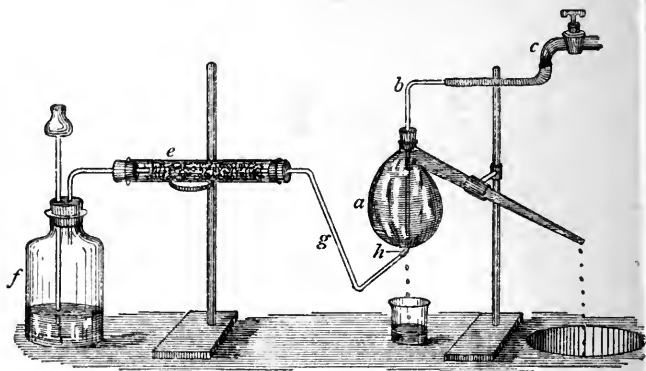


FIG. 108.

a clamp, and connect the tube *b* with a water tap *c*. On turning the tap, water flows into the retort, fills it, and drips out from the neck. Place the retort so that the water drips into a sink.

We are going to burn Inflammable Air. But we must be sure that it is free from moisture. This can be secured by passing it over a substance known as chloride of calcium, which readily takes up moisture. Hence, pack a glass tube *e* with some lumps of chloride of calcium, and supporting it on a ring of a retort-stand or by other means, connect it with a bottle *f* in which Inflammable Air is prepared. Attach to the other end of the tube a glass tube *g* bent as shown in the figure, so that

<sup>1</sup> This form of the experiment is due to Mr. C. M. Stuart; *see Nat.* liii, 347.



the end *h* comes just below the retort. The end *h* of the tube must be narrowed.

As soon as all atmospheric air has been swept out, light the Inflammable Air at *h*. The flame impinges against the cold surface of the retort, and a dew is seen to be formed. Place a small beaker immediately below the retort. Presently a drop of the dew collects and falls into the beaker.

We must obtain sufficient of the dew to enable us to examine its nature. If several students set up similar apparatus, enough will be obtained in half an hour. Otherwise the experiment should be continued for an hour and a half.

**EXPT. 5. To show that the Dew is Water.**—The chemist has tests by which he can recognise very many substances. But water, one of the most common, is not one of the easiest to test for.

Examine the liquid as regards :—

(a) *Smell.*

(b) *Taste.*

(c) Powder some crystals of blue copper sulphate. Heat them gently in a porcelain crucible until they become quite white. Break the white lumps down with a glass rod. Allow them to get quite cold. Add a few drops of the liquid to a portion. The copper sulphate turns blue again. Repeat, using tap-water.

(a) *Boiling point.* Place a few c.c. of the liquid in a test-tube. Hold the bulb of a thermometer just above the liquid by means of a two-holed cork. Heat the liquid carefully in a flame, and observe its boiling point. Repeat, using tap-water.

(e) *Freezing point.* Place a few c.c. in a test-tube. Hold the bulb of a thermometer in the liquid. Freeze the liquid by holding the tube in a freezing mixture. At what temperature does it freeze? Then warm the tube gently; at what temperature does the ice melt? Repeat, using tap-water.

What do you conclude from your observations? The liquid has the properties of water. Hence, it may be concluded that it is water.<sup>1</sup>

Recall Experiment 2. It seems that *water must be a compound of Inflammable Air and oxygen.*

<sup>1</sup> Cavendish proved that under certain conditions a little nitric acid is also formed and can be found in the dew.

In order to make still more sure of this conclusion, Cavendish exploded together large quantities of Inflammable Air and oxygen, and found the dew deposited to be water.

## (B) THE ANALYSIS OF WATER

We have seen how Cavendish proved in 1782 by *synthesis* that water is a compound of Inflammable Air and oxygen. In the following year Lavoisier made this conclusion certain by effecting the *analysis* of water.

The problem to be solved may be stated thus. Given that water is a compound of oxygen and Inflammable Air, how can these be separated from water? Now we know of several

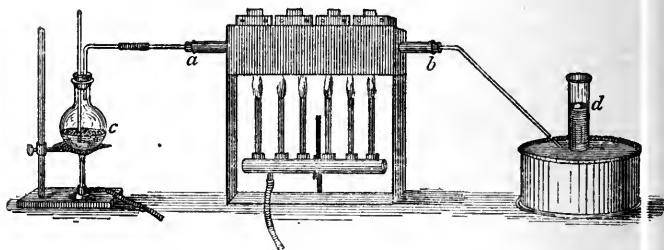


Fig. 109.

substances which combine readily with oxygen when heated, for instance, iron, magnesium, and others. Is it not possible that one of these may combine with the oxygen which we are now assuming to be in water, just as in a former experiment we succeeded in removing oxygen from nitric oxide (Nitrous Air) by passing it over heated metals? If so, Inflammable Air may be set free. Suppose we try this by passing steam over red-hot iron.

**EXPT. 6.<sup>1</sup> To find whether Water is decomposed by Red-hot Iron.**—Pack an iron gas pipe 2 feet in length with clean iron or steel tacks.<sup>2</sup> Place it in the bed of a furnace (Fig. 109). Light the furnace, so that the tube is heated to a *bright red heat*.

<sup>1</sup> Lecture Experiment.—A short "combustion" furnace may be used.

<sup>2</sup> They may be cleaned from grease by heating them in the furnace whilst a current of coal gas is passed over them.

Attach a delivery tube to the end *b* of the iron tube, and have a pneumatic trough and gas jar at hand. Place some water in a 16 oz. flask *c*. Fit it with a cork, through which are passed a straight glass tube dipping below the surface of the water and a bent glass tube, so that it can be connected with the end *a* of the tube in the furnace. Boil the water in the flask. Then connect the flask with the iron tube, so that the steam is passed over the red-hot iron. Place an inverted gas jar *d* full of water over the end of the delivery tube. A gas bubbles up and is collected in the jar. Neglect the first jar which is filled. Fill another. Test the gas with a lighted taper. It burns with a pale flame, so Inflammable Air has been set free, as we hoped. Allow the iron tube to cool, and extract some of the iron nails. Examine them.

It is of interest to note the way in which Lavoisier himself carried out this experiment.

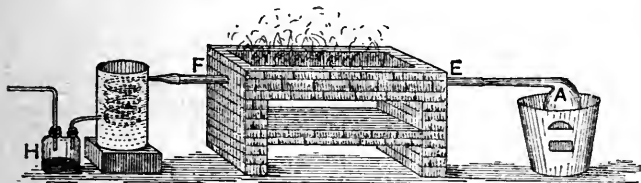


FIG. 110.

**Lavoisier's Experiment to show the Decomposition of Water by Iron.**—A gun barrel *E F* was filled with iron filings, and placed in a furnace. To the end *E* was luted a glass retort *A* containing a known weight of pure water. The end *F* was connected with a worm surrounded by cold water. The other end of the worm passed into a bottle *H*. A delivery tube was led from the flask *H* to a pneumatic trough. The tube *EF* was heated red hot. On lighting a fire under *A*, the water boiled and the steam passed down the gun barrel. The retort was heated until the water had entirely evaporated. Some of the steam passing through the furnace unchanged was condensed to water in the worm, and collected in *H*. The rest was decomposed by the iron, with the result that a gas was

set free. This was collected at the trough, and proved to be Inflammable Air. At the end of the experiment Lavoisier examined the iron in the tube. It had *gained* in weight ; it was "in the state of the black oxide, precisely like that which had been burnt in oxygen." Moreover, Lavoisier observed the quantities obtained ; thus, to quote an experiment,

Weight of water decomposed = weight of water taken in A less weight of water collected in H = 100 grains.

Weight of Inflammable Air produced = 15 grains.

Increase in weight of the iron = 85 grains.

*This was the First Analysis of Water ever made.*—Observe the conclusion to which Lavoisier came. "One hundred grains of water have been decomposed ; 85 of oxygen have combined with iron, converting it into the black oxide, and 15 grains of Inflammable Air have been liberated ; *therefore*, water is composed of oxygen and Inflammable Air in the proportion of 85 parts to 15."

#### PROBLEM

Find whether steam can be decomposed by (a) magnesium, (b) zinc, (c) copper, with liberation of Inflammable Air, employing the apparatus shown in Fig. III :—

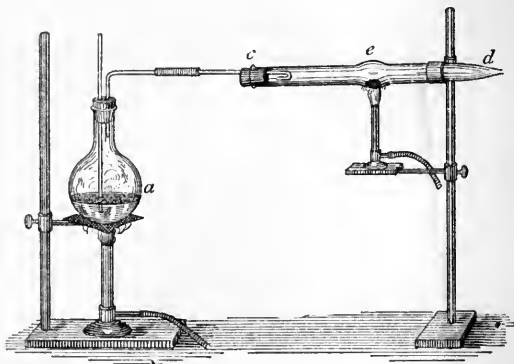


FIG. III.

*a* is a flask in which water is boiled. Through the cork is passed a straight piece of glass tubing dipping below the surface of the water and

a tube bent at right angles. *cd* is a tube of hard glass, on which a bulb *e* has been blown,<sup>1</sup> fixed in a slightly inclined position, and narrowed at the end *d* or fitted with a short straight tube. The cork at *c* is connected with the flask *a* by a piece of glass tubing, one end of which is bent back on itself as shown. Place the metal in *e*. Pass steam through the apparatus, warming the tube *cd* so that the steam may not condense in it, and then heat the bulb *e*.

Find in each case (1) whether Inflammable Air can be kindled at *d*; (2) whether the metal is oxidised.

**A new Name for Inflammable Air.**—Lavoisier, after making this analysis, at once proposed a new name for Inflammable Air. He said: "No name has appeared to us more suitable than that of *hydrogen*, that is to say, 'generative principle of water'; from *hudor*, water, and *geinomai*, I produce." This term was universally adopted, and from this point it will be used in this book.

It is to be noticed that in Lavoisier's analysis of water only *one* constituent was obtained *free*. Cruikshank a few years later first showed that water could be decomposed by the electric current, both constituents being liberated.

**EXPT. 7.<sup>2</sup> To show the Decomposition of Water by Electricity.**—Take a bottle with a wide mouth, and cut off the lower half horizontally, as follows:—Make a scratch on the bottle with a file. Take a pointed piece of glass rod, heat the end red hot, and touch the scratch with it. The glass cracks, and the crack can then be led right round the bottle by means of the hot glass rod, if this is reheated from time to time.<sup>3</sup> Choose a 2-holed rubber cork to fit the mouth, two short jets of glass (wash-bottle tips) to fit the holes in the cork, two platinum wires each 6 in. long, and two pieces of sheet platinum each 2 in. by  $\frac{1}{2}$  in. Put the wires through the glass tubes and seal them in by melting the glass. Fit the glass tubes through the cork. Along the middle line of each platinum sheet prick six holes. Thread the platinum wires in and out through the sheets. Stand the bowl in a tripod-stand and fill it with water (Fig. 112).

<sup>1</sup> Suitable tubes are sold as *reduction tubes*.

<sup>2</sup> Lecture Table Experiment.

<sup>3</sup> Better still, buy ready made a 6 in. glass "dialyser," a kind of basin with a bottle mouth.

Connect the projecting platinum wires by screw clamps to the copper wires from a 4-celled Grove's battery<sup>1</sup> or small dynamo. Nothing is noticed. Add a few drops of dilute sulphuric acid to the water. There is a rapid effervescence from the surface of the platinum sheets. But this soon slackens. Small bubbles have covered the platinum, preventing contact with the water and hindering further action. This could have been avoided by dotting the platinum sheets all over with tiny holes

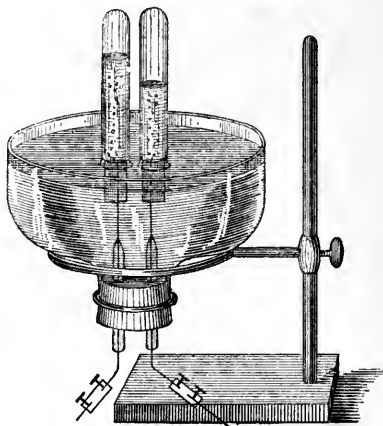


FIG. 112.

pricked by a sharp needle. From the roughened surfaces the bubbles then detach themselves easily.

From what we now know of the composition of water it seems possible that the bubbles may be oxygen or hydrogen. Fill a test-tube with water, invert it over both the platinum sheets and collect the gas. When half full close the end with the thumb, remove the tube, and apply a light. A sharp explosion rings out. The gas is, therefore, probably a mixture of oxygen and hydrogen.

Is the same explosive gas given off by each platinum sheet? Fill two test-tubes with water and invert one over each sheet

<sup>1</sup> See Appendix for Management of Grove's Battery.

Again the current of gas slackens. But raise the tubes so that the platinum sheets are no longer shielded from each other. The bubbles come freely as before, and collect in the tubes. The gas collects much faster in one tube—about twice as fast as in the other. Find with which metal in the battery this is connected. Test the gas in each tube with a burning splinter of wood. The gas which collects slowly quickens the burning of the wood. It is oxygen. The other gas burns with the pale blue flame of hydrogen.

This experiment seems to afford additional evidence that water consists of oxygen and hydrogen, though the student may regard the small amount of sulphuric acid added to the water with suspicion until its part in the change has been clearly proved. In order to complete this experiment it is also clearly desirable to show that nothing else but hydrogen and oxygen results from the decomposition of water by electricity. This was shown after long inquiry to be the case by Sir Humphry Davy.

## CHAPTER XXVI

### THE PROPERTIES OF GASES.—THE DENSITIES OF GASES

#### Solids, Liquids, and Gases

MANY liquids and solids, as we well know, change into gases or vapours when they are sufficiently heated. Water and iodine are readily vaporised, and ether still more so. But the vapours produced are not permanent, and when cooled are condensed or converted into liquids or solids again. We know, however, that there are other substances which are gases at the ordinary temperature. Two hundred years ago, atmospheric air was the only one known. Following the lines of discovery successfully pursued by Priestley, Scheele, Lavoisier, Black and others, we now know of oxygen, nitrogen, hydrogen, carbonic acid gas, nitric oxide and oxide of sulphur, besides common air. Whether these gases will remain gases at all temperatures is so far as we can say at the present point in our enquiry an open question. Since the vapours which can be obtained from liquids are liquefied again by cold, it seems possible, at any rate, that some of these gases might be changed into liquids at a low temperature. We cannot, however, stop to investigate this now.

We are acquainted, then, with three states of matter. But what are the essential differences between solids, liquids, and gases?

**Solids**—RIGID SOLIDS.—Take a piece of steel. It is called a *solid* body. It has a particular shape of its own. Now try to



change its shape by pressing on it with a pair of pincers ; it is not changed in shape perceptibly. And so, in general, a body which is not easily changed in shape when subjected to pressure in a given direction, or to a pull, or to a twist, or to any combination of pressure, pull, and twist (called in general a *stress*), is said to be a *rigid solid*.

If we subject steel to a very great stress in a given direction we find by careful measurement that a change of shape does take place, though the change is small. For instance, a piece of steel pianoforte wire can be extended  $\frac{1}{86}$  of its length by a pull of 106 kilograms weight. Moreover, we find that a *given stress will always produce a definite change in shape* (called in general a *strain*). This last property belongs not only to rigid solids like steel, but to all solids. It is the characteristic property of solids.

**SOFT SOLIDS.**—If we take the pincers and exert upon a piece of india rubber a pressure exactly similar to that which we exerted on the piece of steel, we find that a much greater change of shape is produced. The difference between hard and soft solids lies, therefore, in the ease with which a change of shape may be produced. But with both hard and soft solids a limited pressure produces a limited change of shape.

**ELASTICITY OF SOLIDS WITH REGARD TO CHANGE OF SHAPE.**—If a solid body after being subjected to a stress is relieved from the stress, it may recover either completely or partially its original shape. If it is capable of recovering its shape completely under certain conditions, it is said to be *perfectly elastic with regard to change of shape* under those conditions. For instance, if a weight of 106 kilograms be attached to a suspended steel pianoforte wire its length will be increased by  $\frac{1}{86}$  of the whole ; but as soon as the load is removed the wire will recover its original length. The steel is perfectly elastic under these conditions.

A solid body which does not recover its shape after subjection to a stress is said to be *plastic*. Soap, for instance, is plastic. Soft solid bodies are more often plastic than elastic. India-rubber, however, is both soft and elastic.

**ELASTICITY OF SOLIDS WITH REGARD TO CHANGE OF VOLUME.**—Suppose we subject a solid body, such as a piece of steel or india-rubber, to a uniform squeeze ; that is to say, suppose

we exert equal pressures at the same time on all its surfaces ; we shall tend to diminish its *volume*. In the case of steel a small pressure produces no perceptible diminution in volume, and even the largest pressures we can exert produce so small a diminution in volume that we find it hard to measure it. Such bodies are said to be difficultly compressible, or by an excusable exaggeration *incompressible*. India-rubber, on the other hand, is quite easily compressible. Most solids, whether rigid like steel, or soft like india-rubber, recover their original volume, nearly or entirely, after compression (provided this has not exceeded a certain limit). They are then said to have *elasticity of volume*. A body which recovers completely its original volume after compression under certain conditions is said to be perfectly elastic with regard to change of volume under those conditions.

**Liquids.**—If we melt a piece of ice in a glass beaker we obtain a certain amount of water, which we call a *liquid*. When water is at rest it is usually, as in the present case, partly in contact with a solid body, and the part in contact with the solid body takes the shape of that body. The portion not in contact with a solid body (or with another liquid) is called its *free surface*. When at rest on the earth the greater portion of this free surface is *horizontal* or perpendicular to the direction of a plumb line.

We see at once that a liquid differs considerably in properties from a solid. It has under ordinary circumstances no rigidity at all. It yields indefinitely before the slightest pressure exerted in a single given direction. If we try to squeeze a portion between our pincers it simply *flows* round them.<sup>1</sup> There can, therefore be no question of elasticity with regard to change of shape.

Liquids are almost without exception difficultly compressible or “incompressible.” For example, water is compressed only  $\frac{1}{11.5}$  of its volume under the enormous pressure of 2000 atmospheres. They have, moreover, perfect or almost perfect elasticity of volume.

**Gases.**—Any one of the gases we know, oxygen for instance differs from a solid or a liquid in that it cannot be contained in an open vessel. If we put oxygen into a closed vessel, we

<sup>1</sup> *Mobile* liquids, like water, yield not only indefinitely but almost instantaneously to the slightest pressures. *Viscous* liquids like treacle yield indefinitely but slowly

find that, however big the vessel, it quickly fills it entirely and exerts a pressure on every part of the vessel. And if we were to make a hole anywhere in the vessel the oxygen would escape.

Like liquids, gases have *no rigidity*.

Unlike solids and liquids, all gases are *easily compressible*.

**EXPT. 1. To show that a Gas can change its Volume.**

—Partly fill a bladder, or an india-rubber balloon with air or some other gas. Tie it up. Place it under the receiver of an air-pump, and exhaust the air (Fig. 113). The balloon swells visibly. Admit the air again. It shrinks to its former size. Gases are much more compressible than solids or liquids, and recover their original volumes completely after compression. But we shall study the compressibility of gases more closely presently.



FIG. 113.

Another difference to be noted between gases on the one hand, and solids and liquids on the other, is that gases are as a rule much *less dense* or lighter, bulk for bulk, than liquids and gases. We shall now proceed to consider the densities of gases more fully.<sup>1</sup>

## THE DENSITIES OF GASES

Now that we are acquainted with several gases, which are permanent under ordinary conditions, it will be well to study them more closely, so as to attain a better knowledge of the gaseous state.

We know in a general way that gases differ in density or weight. A soap bubble filled with hydrogen or a balloon full of coal gas will rise in the air. Hydrogen can be poured upwards, and carbonic acid downwards. We have not, however, as yet made any exact measurements of the relative densities of gases.

Now the mass of a unit volume of any substance is called its density. Hence for gases, *Density means the mass of 1 litre of the gas*. For example, experiment has already shown that 1 litre of hydrogen weighs .09 gram. Hence it is said that the density of hydrogen is .09.

<sup>1</sup> NOTE.—In this chapter the densities of gases are discussed apart from the influence of temperature and pressure.

It is often convenient to compare the density of a gas with that of some other well-known gas. *Air* is frequently taken as the standard. Then the *Relative Density of a Gas* means the weight of any volume of the gas divided by the weight of an equal volume of the standard gas, air. For example, if the relative density of carbonic acid gas is stated as  $1\frac{1}{2}$ , it means that it is  $1\frac{1}{2}$  times as heavy as an equal volume of air.

**The Density of Air.**—The first attempt to weigh a gas was made in the case of air, which for a long time was the only gas known. This was tried by Boyle about 1660.

EXPT. 2.<sup>1</sup> **An Attempt to weigh Air.**—Obtain an open gold-beater's skin balloon. Place it along with a piece of thread on one pan of a balance, and weigh it. Attach it to a bellows, inflate it, and tie up the neck with the thread. Weigh again. The weights are the same. Why is this?

Think of the balloon lying on the balance pan, open at the mouth. Air is inside it and around it and the pressure of the air is the same in all parts. Why should any passage of air in or out of the bladder change its weight? If more air is put into the balloon, it does become heavier by the weight of this air. But in air as in water the *upward force on a body immersed in a fluid is equal to the weight of fluid displaced* (see page 54).

If the balloon has a litre more of air put into it then the upward force is increased by the weight of a litre of air. So the weight of air in the balloon is always exactly balanced by the buoyancy of the air around. It is the balloon alone, not its contents, which can be weighed on the balance.

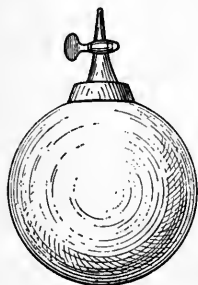


FIG. 114.

**Von Guericke's Method.**—Successful experiments were made about the same time by Otto von Guericke in Magdeburg. He used a large flask with a stopcock (Fig. 114). The capacity of the flask was first determined. The flask was then

emptied of air with the air-pump and weighed. Lastly the air was allowed to enter and the flask was weighed again. For example:—

<sup>1</sup> Lecture Table Experiment.

Capacity of flask ... ..	750 c.c.
Weight when empty ... ..	239.0 gm.
Weight when full of air ... ..	239.9 gm.
Weight of air ... ..	.9 gm.
Then 750 c.c. of air weigh ... ..	.9 gm.
Hence, 1000 c.c. of air weigh... ..	$.9 \times \frac{1000}{750}$ gm.
	= 1.2 gm.

**EXPT. 3.<sup>1</sup> To find the Weight of 1 Litre of Air (Laboratory Method).—**

Fit up an 8 oz. round-bottomed flask or bolt-head with a rubber cork, glass tube, rubber tube and clip, as shown in Fig. 115. Put 50 c.c. water in the flask. Support the flask on a retort-stand over wire gauze. Open the clip. Boil the water. Let it boil steadily for 5 minutes. The air which was in the flask will be swept out by steam. (What will happen if the flame is taken away from below the flask for a few seconds?) Close the clip, quickly remove the flame, thoroughly cool the flask under the tap. (Why does the water boil again?) Dry the outside of the flask. When *cold* weigh the flask by hanging it from the hook of the balance by a wire. Open the clip. Air will hiss in. Weigh again. Find the volume of the air left in the flask, thus :—pour the water left in the flask into a measuring jar ; find the whole volume of the flask ; subtract.

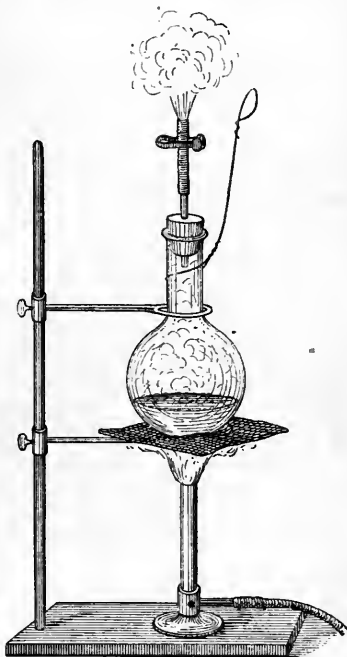


FIG. 115.

<sup>1</sup> Adapted from Ramsay's *Experimental Proofs of Chemical Theory*, p. 26.

Here is the result of an experiment :—

Weight of flask without air ... ..	128.47 gm.
Weight of flask with air ... ..	128.81 gm.
Difference = weight of air... ..	.34 gm.
Volume of water left in flask ... ..	52 c.c.
Volume of whole flask ... ..	340 c.c.
Difference = volume of air ... ..	288 c.c.
Then 288 c.c. air weigh ... ..	.34 gm.
Hence, 1000 c.c. air weigh ... ..	$\frac{1000}{288} \times .34$
	= 1.18 gm.

### **<sup>1</sup> Alleged Inaccuracy of the above Experiments.—**

With an ordinary air-pump (*e.g.*, Tate's) it is difficult to remove more than  $\frac{5}{8}$  of the air from a globe. Mercury pumps (*e.g.*, Sprengel's) will give a much greater exhaustion. Again in the last experiment it is possible that a small amount of the air may not be swept out by the steam. But there is another source of error.

As a cork floats on water or a balloon in the air, so too a glass globe is buoyed up by the air around it. Although the force is not enough to make the globe float, it does make it seem lighter. A globe which displaces a litre of air is buoyed up by a force equal to the weight of a litre of air. We have seen that this is about 1.18 gram for a litre globe. But the brass weights on the other pan are also buoyed up, though to a less extent because they are smaller, say by a force of .004 gram weight. Then on the whole such a glass globe must seem to weigh 1.18 — .004 or 1.176 gram lighter than its weight in vacuo. But this error affects both weighings; and, therefore, as the results are subtracted, the errors seem to cancel each other.

Then are these corrections for the air displaced needless? The numbers 1.18 and .004 depend on the weights of air displaced. Now the weight of the air depends, as we shall presently find, upon its temperature and pressure, and these may change during the course of an experiment. In that case *different* corrections may have to be made out and applied to each separate weighing.

**<sup>1</sup> Regnault's Method.**—This method enables the observer

<sup>1</sup> To be read by advanced students only.

to conduct the experiment so that no corrections need be applied, however much temperature and pressure may vary during the experiment.

Two glass globes are required of as nearly the same capacity as possible. They are suspended by wires at the two ends of the beam of a balance without removing the pans, and they are exactly counterpoised by adding a small piece of glass tubing to the lighter globe (Fig. 116). The left-hand globe is then exhausted

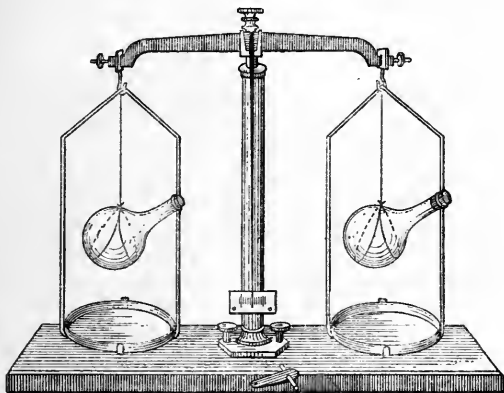


FIG. 116.

as completely as possible by an air-pump, then closed or sealed up, and replaced above the left-hand pan. To restore equilibrium the necessary weights are added to the left-hand pan. The capacity of the left-hand globe must also be determined. For example :—

Capacity of globe ... ..	300 c.c.
Counterpoise added to empty globe	·363 grm.
If 300 c.c. of air weigh... ..	·363 grm.
1000 c.c. of air weigh... ..	$\frac{363}{300}$ grm.
	= 1·21 grm.

In an experiment conducted according to this method, therefore, by employing 2 globes of equal capacity, any change in the

buoyancy of the air affects both ends of the balance beam to the same extent. For the volume of air displaced by the weights used in the second weighing is so small as to be negligible.

This method may be applied to find the density of any gas. We shall, however, proceed to show how the density of carbonic acid may be determined by another method. We know that carbonic acid gas is heavier than air. What, then, does a litre of carbonic acid weigh?

**EXPT. 4. To find the Weight of 1 Litre of Carbonic Acid Gas.**—Carefully clean and dry a 16 oz. flask. Fit it with a good cork. Weigh it accurately. Remove the cork. Pass the gas from a generator containing marble and hydrochloric acid through a drying bottle containing strong sulphuric acid (to dry the gas) and into the flask, filling it by downward displacement. Test with a burning match from time to time, and when the flask is full remove the delivery tube. Cork immediately. Weigh again. Then invert the flask in a trough containing dilute (three per cent.) caustic soda solution. Remove the cork below the surface of the solution, and agitate mouth downwards until all the carbonic acid has been absorbed. Cork it again. The size of a small air bubble remaining shows how nearly all the air is removed. Remove the flask and ascertain the volume of caustic soda in the vessel by pouring it into a measuring jar.

In an experiment,

{	Weight of flask filled with air	... ..	= 66·855 gm.
	"    "    carbonic acid	... ..	= 67·173 gm.
	difference	... ..	= 0·318 gm.
	Volume of carbonic acid...	... ..	= 497 c.c.

Now, weight of 497 c.c. air ... .. =  $1·29 \times \frac{497}{1000}$  or 0·641 gm.

Hence, weight of 497 c.c. carbonic acid = 0·641 + 0·318.  
= 0·959 gm.

And, weight of 1 litre of carbonic acid gas ... =  $0·959 \times \frac{1000}{497}$   
= 1·97 gm.

Do your results confirm this?

**Summary.**—The most accurate measurements of gaseous densities have been made by Regnault's method. Recently, in



the hands of Lord Rayleigh, results of extraordinary accuracy have been obtained.<sup>1</sup> For example, the density of hydrogen is determined as  $\cdot 09001$ , or very nearly  $\cdot 09$ .<sup>2</sup> Hydrogen is found to be very much lighter than all the common gases, and consequently it is very convenient to refer their densities to hydrogen as a standard. In that case their density relative to hydrogen can be stated very approximately as whole numbers, as in the third column of the following table :—

	Density or Weight of 1 Litre.	Relative Density Air = 1.	Relative Density Hydrogen = 1
Hydrogen ... ..	$\cdot 09$	$\cdot 0696$	1
Nitrogen ... ..	1 $\cdot 26$	$\cdot 972$	14
Air ... ..	1 $\cdot 29$	1 $\cdot 000$	14 +
Nitric oxide ... ..	1 $\cdot 35$	1 $\cdot 036$	15
Oxygen ... ..	1 $\cdot 44$	1 $\cdot 105$	16
Carbonic acid ... ..	1 $\cdot 98$	1 $\cdot 529$	22
Sulphur dioxide ... ..	2 $\cdot 88$	2 $\cdot 247$	32

#### FACTS WORTH REMEMBERING :—

1 litre of air weighs 1 $\cdot 29$  gram

1 litre of hydrogen weighs  $\cdot 09$  gram.

1 gram of hydrogen fills  $\frac{1}{\cdot 09}$  or 11 $\cdot 11$  litres.

#### PROBLEMS

1. Use the apparatus employed in Experiment 3, but without water inside. Weigh it dry and full of air. Suck out some of the air. Weigh again. Find the volume of air removed. Hence calculate the weight of 1 litre of air.

2. Devise a method of using the same apparatus to find the weight of 1 litre of coal gas.

<sup>1</sup> One splendid result of Rayleigh's accurate weighings has been his discovery (1894) in conjunction with Professor Ramsay, of another gas in the atmosphere, which is present in extremely small quantity. This gas has been named *argon*.

<sup>2</sup> Professor Morley has (1895) obtained the value  $\cdot 08987$ .

## EXAMPLES XII

## The Densities of Gases

1. A flask weighs 130 grams when full of air, and 129.84 grams when some of the air has been sucked out. When opened under water 125 c.c. of water enter. Find the weight of a litre of air.

2. A flask of 750 c.c. capacity weighs .9 gram more when full of air at 29 inches mercury pressure than when full of air at 2 inches pressure. Find the weight of a litre of air at 30 inches pressure.

3. If a litre of air weighs 1.29 gram, find how many c.c. 1 gram of air will fill.

4. Find the weight of 1 cu. ft. of air in ounces; assuming 1 linear foot = 30 cm. long, 1 oz. = 28 grams, 1 litre air = 1.26 gram.

5.<sup>1</sup> A flask of 250 c.c. volume weighs 125 grams when full of air. What will it weigh when empty?

6.<sup>1</sup> A closed flask of 10 litres external volume is balanced by a brass kilogram weight. Find the weight of the flask in vacuo. Relative density of brass = 8.

7.<sup>1</sup> If 250 c.c. of a gas weigh .09 gram more than 250 c.c. of air, find the weight of 1 litre of the gas.

8.<sup>1</sup> If coal gas is half as heavy as air, find the lifting power of a balloon of 100 litres capacity.

9.<sup>1</sup> If 750 c.c. of a gas weigh .39 gram less than 750 c.c. of air, find the weight of 1 litre of the gas.

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<sup>1</sup> In questions 5 to 9 assume that 1 litre of air weighs 1.2 gram.

## CHAPTER XXVII

### THE PROPERTIES OF GASES—BOYLE'S LAW

WE have already satisfied ourselves that the air, though invisible, has weight. About 230 years ago the Honourable Robert Boyle working in Oxford perfected the air-pump or "pneumatical engine," as he called it, and proceeded to make many interesting experiments. He was soon persuaded that "there is a spring or elastical power in the air we live in." He made the experiment described on p. 231, and saw in the expansion of the bladder as the air was gradually removed from the receiver an evidence of the pressure or force of the "spring" of the air inside the bladder.

Again, he fitted a barometer into the receiver of an air-pump, and found to his delight that as the air pressure upon the mercury in the cistern was exhausted, the space above the mercury in the barometer tube increased at the same time, and the mercury fell.

This experiment suggested to him that it was worth while to investigate what relation there was between the spring or pressure of the air in the barometer tube and the volume it occupied, for, said he, "it is probable that the discovery would not be unuseful." He finally solved this question successfully in a way which we can readily repeat.

**EXPT. 1. To find the Relation between the Volume of Air and the Pressure it exerts.**—Take an ordinary glass tube, about 110 cm. long, and seal one end *a* in a flame. Bend the tube at a point about 20 cm. from the sealed end, until the

two limbs are parallel to one another. Cut a groove in a cork disc. Draw out the tube of a thistle funnel so that it can be inserted in the glass tube at *c*. Support the tube by means of a clamp upon the cork.

Now pour in enough mercury to fill the bend of the tube. Take the tube out of the clamp, and bring the mercury to the same level in both limbs of the tube by inclining the tube

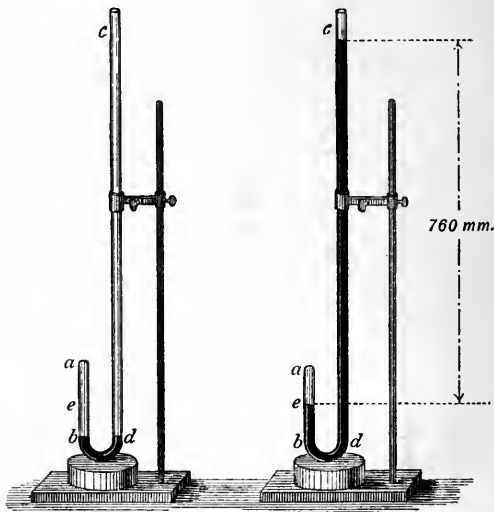


FIG. 117.

frequently, and so permitting air to pass from one limb to the other. Fix the tube in the clamp again. The enclosed air is now at the pressure of the atmosphere outside. Mark the level of the mercury in the closed limb at *b* with a small label. Measure the length *ab* with a metre scale. Gum a small label at a level *e*, half-way between *a* and *b*.

Then carefully pour mercury into *cd* through the funnel. The volume of the air in *ab* is diminished, for the level of the mercury rises gradually up *ba*. Proceed in this way until the mercury has risen exactly half-way up *ba*. Now observe that the mercury

stands at a much higher level in the limb *cd*. Measure the difference in the level of the mercury in the two limbs with the metre scale. Is it not about 760 mm.?

Therefore, the pressure of the enclosed air is now balanced by that of the atmosphere and that due to 760 mm. of mercury. But the pressure of the atmosphere is equal to that of 760 mm. of mercury. Therefore the pressure of the enclosed gas is *double* the atmospheric pressure, while the gas occupies *half* the original volume.

Boyle's account (1662) of his own experiment is worth reading, and after performing the above experiment it will be easier to follow the rather involved sentences in which it is described.<sup>1</sup>

"We took then a long glass tube, which, by a dexterous hand and the help of a lamp, was in such a manner crooked at the bottom, that the part turned up was almost parallel to the rest of the tube, and the orifice of this shorter leg of the syphon (if I may so call the whole instrument) being hermetically sealed, the length of it was divided into inches (each of which was subdivided into eight parts) by a streight list of paper, which containing those divisions, was carefully pasted all along it. Then putting in as much quicksilver as served to fill the arch or bended part of the syphon, that the mercury standing in a level might reach in the one leg to the bottom of the divided paper, and just to the same height or horizontal line in the other; we took care by frequently inclining the tube, so that the air might freely pass from one leg into the other by the sides of the mercury (we took, I say, care), that the air at last included in the shorter cylinder should be of the same laxity with the rest of the air about it. This done we began to pour quicksilver into the longer leg of the syphon, which by its weight pressing up that in the shorter leg did by degrees streighten the included air; and continuing this pouring in of quicksilver till the air in the shorter leg was by condensation reduced to take up but half the space it possessed (I say, possessed, not filled) before; we cast our eyes upon the longer leg of the glass, on which was likewise pasted a list of paper carefully divided into inches and parts, and we observed not without delight and satisfaction that the quicksilver in that longer part of that tube was 29 inches higher than the other. . . . Here the same air being brought

<sup>1</sup> Boyle's Complete Works, vol. 1, p. 100.

to a degree of density about twice as great as that it had before, obtains a spring twice as strong as formerly."

### PROBLEMS

Making use of the U-tube used in Experiment 1, find the change in pressure which occurs when—

- the volume is reduced to three-quarters of the original volume.
- the volume is reduced to two-thirds of the original volume.
- the volume is reduced to half of the original volume.

The result of all these experiments may be summed up in the following statement, which in honour of Boyle is usually known as

**Boyle's Law.**—*When the temperature remains constant, the volume occupied by a definite mass of air varies inversely as its pressure, e.g. if the pressure is increased 2, 3, 4, 5 times, the volume will be reduced to  $\frac{1}{2}$ ,  $\frac{1}{3}$ ,  $\frac{1}{4}$ ,  $\frac{1}{5}$ , of the original volume.*<sup>1</sup>

The apparatus we have used is suitable for studying the change in the volume of air when the pressure is increased above that of the atmosphere. It is not suitable for finding whether Boyle's law applies also when the pressure is reduced; yet it is most desirable to learn whether this is so or not.

**EXPT. 2. To test Boyle's Law for Pressures less than that of the Atmosphere.**—Take a straight tube A F (Fig. 118), about 3 feet long, sealed at one end, like the one used in Experiment 4, p. 60. Fill rather more than half the tube with mercury. Measure the length A B of the column

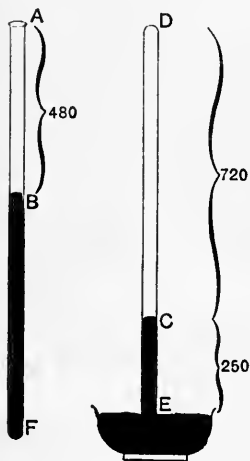


FIG. 118.

of air still in the tube say 480 mm. Hold the thumb firmly over the open end A. Invert the tube in a basin of mercury. The air rises to the top. Remove the thumb. The mercury

<sup>1</sup> Boyle must share the honour of arriving at this important generalisation with Mariotte, whose treatise *De la Nature de l'Air*, was published in Paris in 1676.

drops and the air expands. Measure the length D C of the air space now, say 720 mm. Also measure the length CE of the mercury column, say 250 mm. Then,

the volume of the air is  $\frac{720}{480}$  or  $\frac{3}{2}$  of what it was.

What is now the pressure of the air in D C? The pressure of the air in A B was that of the atmosphere. Read the barometer, say 750 mm. Now,

pressure of the air in D C +  
pressure of the mercury column C E } balance the external pressure of  
the atmosphere.

Hence,

Pressure of the air in D C = 750 - 250 or 500 mm. of mercury.

Hence, the pressure of the air is  $= \frac{500}{750}$  or  $\frac{2}{3}$  of what it was.

Now  $\frac{3}{2}$  is the inverse ratio of  $\frac{2}{3}$ , and consequently the volume of the air has changed in the inverse ratio of the change in the pressure.

Boyle was acquainted with but one gas—viz., atmospheric air. We now know other gases, and therefore it is important to try whether Boyle's law is true for them also.

**EXPT. 3. To find whether Boyle's Law is true for other Gases, for example, Carbonic Acid Gas.**<sup>1</sup>—Fit a small strong wide-necked bottle F with a 3-holed stopper (Fig. 119). Introduce through two of the holes two glass tubes A and B each one metre long, which are drawn out near their upper ends. Fix through the third hole a tube C open at both ends. Obtain also a long glass rod which can slide freely within C. Fill the bottle with mercury. Fit the cork firmly in the neck, and bind it down with wire. Place the rod in C.

Pour mercury into the tube C until the two other tubes are full of mercury. Connect A to a generator of carbonic acid gas. Leave B open to the air. Now raise the rod slowly. The mercury falls in A and B, and car-



FIG. 119.

<sup>1</sup> Lecture Table Experiment.

bonic acid is drawn into A, and air into B. Manipulate the supply of carbonic acid so that the mercury levels E and D are about halfway down the tubes A and B, and the same in each. Then seal the tubes at A and B.

Now lower the plunger in C. The level of the mercury in C is raised, and therefore the pressure upon the gases in A and B is increased, and to the same extent. Notice that the two surfaces E and D move up together, and are always on the same level. This shows that under the same pressure the gases are equally compressible. Also raise the plunger. The surfaces E and D sink together.

It is evident that under changing pressure the volume occupied by the carbonic acid changes in precisely the same way as that of the air. We are, therefore, justified in concluding that Boyle's law applies to carbonic acid as well as to air. In the same way it can be proved that the law is equally applicable to any of the other permanent gases with which we are acquainted.

We may therefore re-state Boyle's Law thus :—

*When the temperature remains constant the volume occupied by a definite mass of any gas varies inversely as its pressure.*

Our experiments have only been made over a small range of pressure. We have therefore no right to assume that if the pressure on a given mass of gas is increased 100 times, its volume will be reduced exactly 100 times. All we can rightly conclude is that Boyle's Law holds for gases within the range of pressure covered by our experiments and with the degree of accuracy of measurement possible in our observations.

#### PROBLEM

Fit up the apparatus shown in Fig. 120, using a-litre flask, rubber stopper, glass T-piece, rubber tubing, two clips, glass tubes, and basin of mercury. Open the clips. Suck at A. Mercury rises up B. Close the clips. Find whether the air in C has expanded according to Boyle's Law.<sup>1</sup>

**Mathematical Expression of Boyle's Law.**—Let a certain mass of gas have a volume  $V_1$  and a pressure  $P_1$ . Let its volume be changed to  $V_2$ , when its pressure becomes  $P_2$ .

<sup>1</sup> This problem is taken from Ramsay's *Experimental Proofs of Chemical Theory*, p. 23, where it is worked out in detail.



Then by Boyle's Law,

$$\frac{V_1}{V_2} = \frac{P_2}{P_1} \quad . \quad . \quad . \quad . \quad . \quad (1).$$

or

$$V_1 P_1 = V_2 P_2 \quad . \quad . \quad . \quad . \quad . \quad (2).$$

Similarly, if it be made to take other volumes  $V_3$ ,  $V_4$ , &c.,

$$V_1 P_1 = V_3 P_3 = V_4 P_4, \text{ \&c.}$$

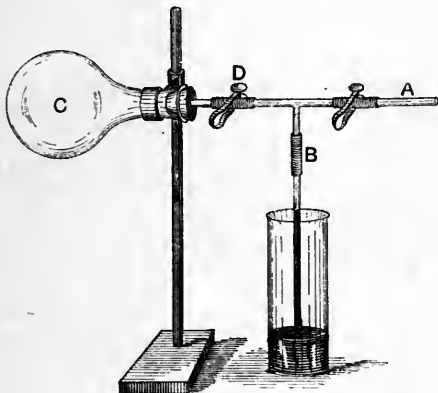


FIG. 120.

The product of pressure and volume, therefore, remains constant, or

$$P V = C \text{ (a constant)} \quad . \quad . \quad . \quad (3).$$

From (2),

$$\frac{V_2 P_2}{V_1 P_1} = 1 \quad . \quad . \quad . \quad . \quad (4).$$

This form of the equation is useful in testing the results of experiments.

EXAMPLE.—In an experiment  $V_1 = 338.5$  c.c.,  $P_1 = 595.2$  mm.,  $V_2 = 274.5$  c.c.,  $P_2 = 736.2$  mm.

Then,

$$\begin{aligned} \frac{V_2 P_2}{V_1 P_1} &= \frac{2745 \times 7362}{3385 \times 5952} \\ &= 1.003 + \end{aligned}$$

Hence the above data agree very nearly with Boyle's Law.

Again, from equation (2)

$$V_2 = V_1 \times \frac{P_1}{P_2} \quad (5).$$

This equation is useful in calculations like the following :—

EXAMPLE.—What volume will be occupied by 250 c.c. of air at 720 mm. when the pressure changes to 750 mm. ?

Here,

$$V_1 = 250 \text{ c.c.}, P_1 = 720 \text{ mm.}, P_2 = 750 \text{ mm.}$$

Hence,

$$\begin{aligned} V_2 &= 250 \times \frac{720}{750} \text{ c. c.} \\ &= 240 \text{ c.c.} \end{aligned}$$

**The Measurement of Gases collected over Water or Mercury.**—When the volume of a gas standing over water is measured, the liquid levels should be the same within and without the tube. This can generally be arranged for. But if not, a calculation must be made.

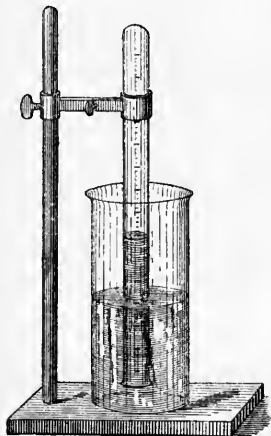


FIG. 121.

EXAMPLE.—100 c.c. of air are collected in a tube over water (Fig. 121), when the barometer is at 750 mm., the water standing 27 cm. higher inside the tube than without. What volume would the air occupy at the atmospheric pressure 750 mm. ?

Now, since the relative density of mercury is  $13\frac{1}{2}$ , 27 cm. of water produce the same pressure as

$$\frac{270}{13\frac{1}{2}} \text{ mm. or } 20 \text{ mm. of mercury.}$$

Then, the external air pressure of 750 mm. is balanced by the internal air pressure, together with that of 20 mm. of mercury.

Hence, the internal air pressure = 750 - 20 mm.

$$= 730 \text{ mm.}$$

Here,  $V_1 = 100 \text{ c.c.}, P_1 = 730 \text{ mm.}, P_2 = 750 \text{ mm.}$

Hence,

$$\begin{aligned} V_2 &= 100 \times \frac{730}{750} \\ &= 97.3 \text{ c.c.} \end{aligned}$$

**Graphical Expression of Boyle's Law.**—Draw two lines OX and OY at right-angles (Fig. 122). Measure distances in any units, say inches, along each. Let distances along OX represent volumes, and distances along OY pressures.

The point A is distant 1 inch from OX and 1 inch from OY, and may be taken to represent a given mass of some gas at unit pressure and unit volume. Similarly, the point B represents the same mass of gas at double volume and half the pressure. Again, C represents the gas at half the volume and double the pressure. Notice that in each case the rectangle enclosed by

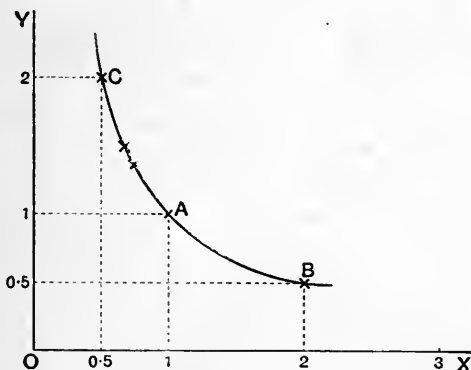


FIG. 122.

the dotted lines, representing volumes and pressures, has an area of 1 square inch. The product, then, of pressure and volume is represented by a constant area (cf. equation 3, p. 245).

Find other points representing the same mass of gas when its volume is  $\frac{2}{3}$ ,  $\frac{3}{4}$ ,  $1\frac{1}{2}$ , &c. Notice that these points lie on a curve. Draw the curve C A B. This line C A B is the graphic representation of Boyle's Law. The curve is what mathematicians call a *rectangular hyperbola*.

Obtain some paper ruled in squares, say twelve squares to the inch, and draw for yourself a "rectangular hyperbola."

## EXAMPLES XIII

## Boyle's Law

1. Given  $V_1 = 99$  c.c.,  $P_1 = 750$  mm.,  $V_2 = 101$  c.c., find  $P_2$ .
2. 1,440 c.c. of air were enclosed in a bell-jar over water on a day when the barometer showed 750 mm. What volume would the air occupy if the barometer had fallen to 720 mm. ?
3. 100 c.c. of air were measured at 750 mm. The volume afterwards changed to 101 c.c. Find the new pressure.
4. Find the capacity in cubic inches of a steel oxygen cylinder which contains under a pressure of 1,800 lbs. per square inch enough oxygen to fill 20 cubic feet under a pressure of 15 lbs. per sq. in. (assuming Boyle's Law to hold at the high pressure).
5. Find the pressure in lbs. per sq. in. required to force 40 cu. ft. of coal gas at 15 lbs. per sq. in. into a space of  $\frac{1}{2}$  cu. ft.
6. By how much must the pressure on 1,000 c.c. of air at 760 mm. be increased to make the volume 50 c.c. less.
7. What change will there be in the volume of 1,000 c.c. of air at 30 ins. of mercury pressure when the barometer falls to  $28\frac{3}{4}$  ins. ?
8. In a Boyle's Law Experiment with a syphon tube there was originally 6 ins. length of air in the shorter limb. The barometer stood at 30 ins. Mercury was poured into the open limb until it was 15 ins. higher there. Find the length of air in the inner tube (*see* Fig. 117, p. 240).
9. In another similar experiment, 12 ins. length of air was enclosed in the shorter limb. What height of mercury would compress the air to 4 ins. length ?
10. Two flasks of 500 c.c. and 1,000 c.c. are connected by a stop-cock. The 1,000 c.c. flask is vacuum. The 500 c.c. flask contains air at 750 mm. pressure. Find the pressure after opening the tap.
11. Are the following figures in agreement with Boyle's Law? Find  $\frac{V_2 P_2}{V_1 P_1}$  to 3 places of decimals.  
 $V_1 = 250$  c.c.,  $P_1 = 760$  mm.,  $V_2 = 253$  c.c.,  $P_2 = 750$  mm.
12.  $V_1 = 240$  cu. ins.,  $P_1 = 29$  ins.,  $V_2 = 232$  cu. ins.,  $P_2 = 30$  ins.

## CHAPTER XXVIII

### THE PROPERTIES OF GASES.—THE EXPANSION OF GASES BY HEAT

IT has already been shown (p. 38) that air expands when heated, and contracts on cooling. We have, however, made no endeavour to measure the amount of this expansion, nor have we tried to compare the expansions of different gases.

*John Dalton* made, in 1801, in Manchester the first really important research on this subject, and shortly afterwards *Charles* in Paris arrived quite independently at the same conclusions as Dalton.

Suppose we study in the first instance the expansion of air when heated.

**EXPT. 1.<sup>1</sup> To find the Expansion of 1 Litre of Air when heated 1° C.**  
Dry thoroughly an 8 oz. round-bottomed flask or bolt-head.

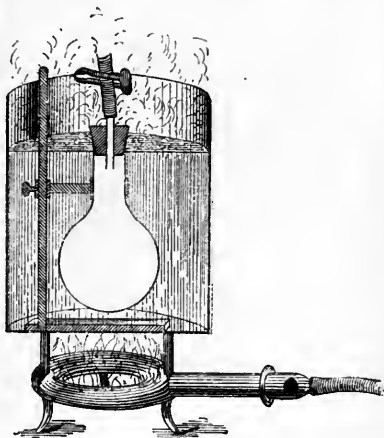


FIG. 123.

<sup>1</sup> Adapted from Ramsay's *Experimental Proofs*, p. 21.

Choose a closely-fitting one-holed rubber stopper. Fit it with a short glass tube, rubber tube and clip<sup>1</sup> as shown in Fig. 123. The tube must fit very tightly into the stopper. Place the stopper firmly in the flask. Fix the flask in a small retort-stand which can be placed within a cylindrical tin vessel. Half fill the tin with water, and heat it over a rose-burner. Open the clip on the flask. As soon as the water begins to boil, place the stand with the flask in the tin so that the flask is immersed in the water up to the stopper. After a minute or two, close the clip. Heat the flask thus for at least 5 minutes, but open the clip for a moment from time to time. Take the temperature of the water. Then tighten the clip, and remove the flask. Cool the flask. Invert the flask in cold water, and open the clip under the water. Water enters the flask. Why? Shake. This will ensure the cooling of the air and flask to the temperature of the water. Lower the flask till the level of the water is the same within and without. Then close the clip under the water. Take the temperature of the water in the flask. Measure the volume of water which has entered the flask. Find also the volume of the whole flask up to the lower surface of the stopper.

The method of calculation will perhaps be best understood by considering first an experiment in which the flask was opened under *ice-cold* water :—

Temperature of boiling water	...	...	...	100° C.
Temperature of cold water	...	...	...	0° C.
Volume of whole flask	...	...	...	360 c.c.
Volume of water entering flask	...	...	...	96 c.c.

Then,

360 c.c. of air contracted 96 c.c. in cooling through 100° C.

Therefore,

(360 - 96) c.c. or 264 c.c. at 0° would expand 96 c.c. when heated to 100° C.

Hence,	1 c.c.	"	"	$\frac{96}{264}$	"	"	100° C.
	1 c.c.	"	"	$\frac{96}{264} \times \frac{1}{100}$	"	"	1° C.
	1000 c.c.	"	"	$\frac{96}{264} \times \frac{1000}{100}$	"	"	1° C.
Or	1000 c.c.	"	"	3.63 c.c.	"	"	1° C.

<sup>1</sup> A stopper of glass rod may be used in place of the clip.

*Secondly*, consider an experiment in which the flask was opened under water at the *ordinary temperature* :—

Temperature of hot water	...	...	...	100° C.
Temperature of cold water	...	...	...	16° C.
Volume of whole flask	...	...	...	360 c.c.
Volume of water entering flask	...	...	...	85 c.c.

Then,

360 c.c. of air in cooling from 100° to 16° (*i.e.* through 84°) contracted 85 c.c.  
*i.e.*, 360 c.c. in cooling through 84° contracted ... .. 85 c.c.

Hence,

360 c.c. in cooling through	1°	would contract	...	...	...	...	$\frac{85}{84}$ c.c.
360 c.c. " "	100°	"	...	...	...	...	$\frac{85}{84} \times 100$ c.c.

Or,

360 c.c. would contract 101·2 c.c. in cooling through ... .. 100° C.

Therefore,

(360 - 101·2) c.c. or 258·8 c.c. at 0° would expand 101·2 c.c. when heated 100° C.

Hence, 1 c.c. " "  $\frac{101\cdot2}{258\cdot8}$  c.c. " 100° C.

1000 c.c. " "  $\frac{101\cdot2}{258\cdot8} \times \frac{1000}{100}$  c.c. " 1° C.

Or, 1000 c.c. " " 3·91 c.c. " 1° C.

More accurate experiments than can be made with this simple apparatus show that the expansion of 1 litre of air at 0° when heated to 1° C is 3·66 c.c. Dalton himself obtained the value 3·63.

Notice that we have assumed in these calculations that the expansion during 1 degree is the same wherever that degree is upon the scale of temperature. By repeating the experiment through different ranges of temperature, say 100° to 0°, 100° to 30°, 100° to 50°, we could find whether we are led in each case to the same result. Careful experiment proves this to be so, and therefore the assumption made is justified.

Suppose we have equal volumes of *different* gases, will they expand to the same extent as air, when heated through the same range of temperature?

**EXPT. 2.<sup>1</sup> To compare the Expansion of Carbonic Acid Gas with that of Air.**—Fit up the apparatus shown in

<sup>1</sup> Lecture Table Experiment.

Fig. 124. There are two similar tubes A and B within a wider one. A contains dry air, B an equal length of dry carbonic acid gas. They dip into and are partly filled with dry mercury. The mercury is contained in a stout bottle C. The cork of this bottle is not air-tight. Surrounding the parallel tubes is a steam jacket D. The mercury levels stand side by side at E. Connect the tube F with a flask or tin can in which water is being vigorously boiled. The gases within the tube are heated by the steam. Observe that they both move down to the same level G. This shows that the expansion of the carbonic acid is equal to that of the air.

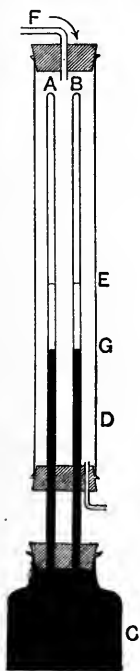


FIG. 124.

**CHARLES'S LAW.**—In the same way as in Experiment 2 the expansion of other gases may be compared with that of air, and the conclusion arrived at may be expressed in words used by Dalton in 1801 when he read his paper on this subject in Manchester.

“The results of several experiments made upon hydrogenous, oxygenous, carbonic acid gas, and nitrous gas, which were all the kinds I tried, agreed with those of common air. . . . Upon the whole, therefore, I see no sufficient reason why we may not conclude that all elastic fluids under the same pressure expand equally by heat.”

It was long forgotten that Dalton had first arrived at this important generalisation. Consequently it is after the Frenchman, who also independently arrived at it, that it is usually known under the name of *Charles's Law*. Viz. :—

*Equal volumes of all gases under the same pressure expand equally for equal increments of temperature.*

Accurate experiments show that 1 litre of all permanent gases at  $0^{\circ}$  will expand the same amount, viz., 3.66 c.c. for every rise of  $1^{\circ}$  C. Hence all gases expand  $\frac{3.66}{1000}$  or  $\frac{1}{273}$  of their volume at  $0^{\circ}$  for every rise of  $1^{\circ}$  C. This number is called their *coefficient of expansion*.



For example,

273 c.c. at $0^{\circ}$	become	273 + 1 c.c. at $1^{\circ}$
„		273 + 2 c.c. at $2^{\circ}$
„		273 + 3 c.c. at $3^{\circ}$
„		273 + 10 c.c. at $10^{\circ}$
„		273 + t c.c. at $t^{\circ}$

We can, therefore, construct a Table such as the following, expressing the volumes of a gas at different temperatures which correspond to 273 volumes at  $0^{\circ}$  C. ; *e.g.*—

293	volumes at $20^{\circ}$ C.
290	„ $17^{\circ}$ C.
288	„ $15^{\circ}$ C.
283	„ $10^{\circ}$ C.
274	„ $1^{\circ}$ C.
270	„ $-3^{\circ}$ C.
263	„ $-10^{\circ}$ C.

**Absolute Zero.**—Then, if on cooling a gas from  $0^{\circ}$  downwards, it contracts  $\frac{1}{273}$  of its volume for every degree, we shall at last reach that temperature at which its volume entirely disappears. This will evidently be at  $-273^{\circ}$ . *This temperature is called the absolute zero*, and it is frequently used as a convenient point from which to measure temperatures.

As a matter of fact, no one has yet succeeded in obtaining so low a temperature ; and most gases are liquefied above  $273^{\circ}$ . Still, *at ordinary temperatures a gas behaves as though its volume would vanish at  $-273^{\circ}$  C.*

*Temperatures reckoned in Centigrade Degrees from Absolute Zero are called Absolute Temperatures A, e.g.*—

$0^{\circ}$ C.	corresponds to	$273^{\circ}$ A.
$20^{\circ}$ C.	„	$273 + 20$ or $293^{\circ}$ A.
$15^{\circ}$ C.	„	$273 - 15$ or $258^{\circ}$ A.
$t^{\circ}$ C.	„	$273 + t$ or $T^{\circ}$ .

If a gas has a volume 1 at  $0^{\circ}$  C. its volume at  $10^{\circ}$  C. will be  $1 \frac{10}{273}$

*i.e.*, the volumes are as  $1 : 1 \frac{10}{273}$

or,  $\frac{273}{273} : \frac{283}{273}$

or,

$273 : 283$

or, as their absolute temperatures,



EXPT. 3.<sup>1</sup> To show that when a Gas is heated at Constant Volume the Pressure is Increased.—Place a small piece of sheet india-rubber at the bottom of an 8 oz. round-bottomed flask. Introduce a test-tube into it so that it rests on the rubber and is supported by the neck of the flask. (Take the temperature of the air in the flask.) Close the flask with a tightly-fitting rubber stopper, through which is passed a long tightly-

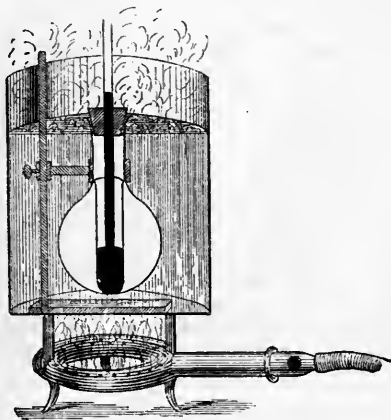


FIG. 125.

fitting narrow glass tube (Fig. 125). Pour some mercury into the tube and see that it takes the same level within and without the tube. Measure the position of this level. Tighten the stopper. Surround the flask with boiling water and steam, as in Experiment 1. The mercury rises in the tube. Take the temperature of the water. Mark the level reached by the Mercury in the tube. Measure the pressure shown by the mercury. For instance, if

Barometer stood at	...	...	...	...	750 mm.
Rise of mercury	...	...	...	...	240 mm.
Original temperature of the air	...	...	...	...	10° C.
Temperature of the boiling water	...	...	...	...	100° C.

<sup>1</sup> Lecture Table Experiment.

Then, for  $90^\circ$  rise of temperature, pressure rose 240 mm.

Hence, ,,  $1^\circ$  ,, ,, would rise  $\frac{240}{90}$  mm.

,,  $10^\circ$  ,, ,, ,,  $\frac{2400}{90}$  or  $26\frac{2}{3}$  mm.

and at  $0^\circ$ , pressure would be  $(750 - 26\frac{2}{3})$  mm.  
 $= 723\frac{1}{3}$  mm.

Therefore,

$$\begin{aligned}\text{coefficient of the increase of pressure for } 1^\circ &= \frac{240}{723\frac{1}{3}} \\ &= \frac{1}{261} +\end{aligned}$$

Thoroughly careful experiments show that when a gas is heated at constant volume the coefficient of the increase of pressure for  $1^\circ$  is  $\frac{1}{273}$ . Observe that this is exactly the same as the coefficient of expansion when a gas is heated at constant pressure. Can you explain this apparently remarkable coincidence? Hence, *when a gas is heated at constant volume the pressure is proportional to the absolute temperature.*

### Mathematical Expression for a Gas under different Conditions of Pressure, Volume and Temperature.—

Let us suppose that a gas in the state denoted by  $P_1, V_1, T_1$ , is converted into the state denoted by  $P_2, V_2, T_2$ , in *two* stages.

(1) While the temperature remains constant at  $T_1$ , change  $P_1$  to  $P_2$ .

Let  $V_1$  in consequence become  $V$ .

Then,

$$P_1 V_1 = P_2 V \quad (\text{Boyle's Law}),$$

$$\text{or} \quad V = \frac{P_1 V_1}{P_2}$$

(2) Then while the pressure remains constant at  $P_2$ , change  $T_1$  to  $T_2$ .

Let  $V$  in consequence become  $V_2$ .

Then,

$$V T_2 = V_2 T_1 \quad (\text{Charles's Law}),$$

But

$$V = \frac{P_1 V_1}{P_2}$$

Hence,

$$\frac{P_1 V_1}{P_2} \times T_2 = V_2 T_1$$

or,

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \dots \quad (1).$$

which is the expression required.

Observe that if we put  $V_1 = V_2$  in equation (1), we have

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

or, when the volume remains constant, the pressure is proportional to the absolute temperature, as we have already shown by experiment to be the case.

EXAMPLE.—819 c.c. of air are measured at  $13^\circ$  C. and 741 mm. What would the volume be at  $0^\circ$  C. and 760 mm.?

Since  $0^\circ$  is a lower temperature than  $13^\circ$ , the volume will be less at  $0^\circ$ , and the volume corrected for temperature will be  $819 \times \frac{273}{286}$

Again, since 760 mm. is a greater pressure than 741 mm., the volume will be less at 760 mm. in the ratio  $\frac{741}{760}$

Hence,

$$\begin{aligned} \text{volume at } 0^\circ \text{ and 760 mm.} &= 819 \times \frac{273}{286} \times \frac{741}{760} \\ &= 762.3 \text{ c.c.} \end{aligned}$$

N.B.—Gas volumes are generally corrected to  $0^\circ$  C. and 760 mm. of mercury.

N.T.P. means normal Temperature and Pressure; i.e.,  $0^\circ$  C. and 760 mm. of mercury.

## EXAMPLES XIV

### Charles's Law

1. Find from each of the following the temperature at which the volume of the air would vanish, if it went on contracting at the same rate (*see* Experiment 1)—

	(1)	(2)
Temp. of air when hot	... $100^\circ$ C.	$100^\circ$ C.
Temp. of air when cold	... $0^\circ$ C.	$20^\circ$ C.
Vol. of whole flask	... 414 c.c.	339 c.c.
Vol. of water entering	... 111 c.c.	73 c.c.
		S

2. Find the coefficient of expansion of air from each of the following—

	(1)	(2)
Temp. of air when hot ...	80° C.	99° C.
Temp. of air when cold ...	0° C.	16° C.
Vol. of whole flask ...	441 c.c.	338 c.c.
Vol. of water entering ...	100 c.c.	76 c.c.

3. A litre of air is measured at 13° C., what will be its volume at 0° C. ? (Answer in c.c.)

4. What volume of air at 21° C. will fill one litre at 0° C. ?

5. If a litre of air at 0° C. weighs 1.29 grams, what will be the weight of a litre at 28° C. ?

6. A fire balloon contains 1,000 litres of air at 114° C. How much lighter is this air than 1,000 litres at 0° C.

7. To what temperature must oxygen (density 16) be heated in order to make it as light as hydrogen is at 0° ?

8. At what temperature will air have half the density that it has at 23° C. ?

### EXAMPLES XV

#### Charles's and Boyle's Laws

1. Find the volume at N.T.P. of—

- (a) 100 c.c. of oxygen at 63° C. and 741 mm.
- (b) 100           ,,       77° C. and 800 mm.
- (c) 100           ,,       - 52° C. and 676 mm.

2. A quantity of air measures 500 c.c. at 0° C. and 760 mm.

- (a) Find the volume at 15° C. and 720 mm.
- (b)       ,,       ,,       33° C. and 760 mm.
- (c)       ,,       ,,       52° C. and 950 mm.

3. A quantity of nitrogen measures 20 c.c. at 17° C. and 740 mm.

- (a) Find the temperature if the volume is 40 c.c. at 760 mm.
- (b)       ,,       ,,       ,,       37 c.c. at 600 mm.
- (c)       ,,       ,,       ,,       20 c.c. at 814 mm.

4. A quantity of carbonic acid measures 40 c.c. at 27° C. and 780 mm.

- (a) Find the pressure if the volume is 40 c.c. at 0° C.
- (b)       ,,       ,,       ,,       50 c.c. at 15° C.
- (c)       ,,       ,,       ,,       30 c.c. at 6° C.

## CHAPTER XXIX

### THE PROPERTIES OF GASES—THE PRESSURE IN A MIXTURE OF GASES

**The Problem to be solved.**—If two separate quantities (say 25 c.c. and 50 c.c.) of the same gas (say hydrogen) at the atmospheric pressure are introduced into the same empty space (say 100 c.c.), what will be the pressure of the mixture? By Boyle's Law the 50 c.c. alone would produce a pressure of  $\frac{50}{100}$  or  $\frac{1}{2}$  atmosphere, and the 25 c.c. alone would produce a pressure of  $\frac{1}{4}$  atmosphere. Both together having a total volume of 75 c.c. will produce a pressure of  $\frac{3}{4}$  atmosphere.

But what will happen if the 50 c.c. and the 25 c.c. are volumes of different gases, say oxygen gas and hydrogen? Each separately will obey Boyle's Law. But the mixture? This is a problem which was attacked and solved by John Dalton in 1801, and we also must now endeavour to work it out for ourselves.

The following preliminary question must first be considered. *How is the volume of a mixture of gases related to the volumes of the separate gases?* For example, do 50 c.c. of oxygen and 25 c.c. of hydrogen make a total volume of 75 c.c. when mixed? The answer must be sought by experiment.

**EXPT. 1.<sup>1</sup> To find the Total Volume of 50 c.c. of Oxygen and 25 c.c. of Hydrogen.**—Take a Hempel burette (Fig. 126). This consists of a graduated glass tube *a*, connected at its lower end by a rubber tube with another glass

<sup>1</sup> Lecture Table Experiment.

tube *b*, open at both ends. The tube *a* is narrowed at its upper end, and to it a bent piece of glass tubing *d* can be connected by a short piece of rubber tubing, on which is placed a clip *c*. Open the clip *c*, and pour water into *b* until it flows out from *d*;

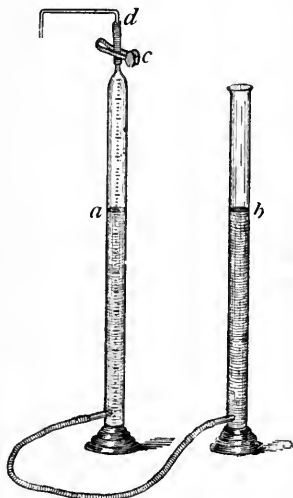


FIG. 126.

then close the clip, and pour a little water out of the tube *b*. Connect *d* with a supply of oxygen gas, open the clip *c*, allow about 50 c.c. to enter and close the clip. Hold the two tubes so that the water level is the same in each. Read the volume. Then pass in on the top of the oxygen about 25 c.c. of the lighter hydrogen gas. Adjust the water level and read the total volume, say 75.2 c.c. Place a clamp on the rubber tube connecting the two tubes. Invert the tube containing the gases once or twice. Unscrew the clamp. Adjust the water level and read the volume again. The change, if any, is too small to be noticed.

We may conclude, then, that the *volume of the mixture of gases is the sum of the volumes of the separate gases*. This can be shown to be so in other cases. In other words, gases mix without change in volume (when the pressure remains unchanged).

We can now proceed to find the relation between the pressure of a mixture of gases and that exerted by each gas.

**DEFINITION OF PARTIAL PRESSURE.**—*The partial pressure of one gas in a mixture of gases is the pressure that it would exert if it alone filled the whole space occupied by the mixture.*

**EXPT. 2.<sup>1</sup> To compare the Pressure of a Mixture of Gases with their Partial Pressures.**—Take a 1,000 c.c. flask. Fit it with a one-holed rubber cork. In the cork, A, fit the

<sup>1</sup> Lecture Table Experiment.



horizontal limb of a T tube (Fig. 127). Support the flask with its neck horizontal in a retort-stand. To the vertical limb of the T tube attach by rubber tubing a glass tube B, 1 foot long. Let this glass tube dip down into a tall jar C, containing mercury. Fix a half metre scale into this jar. To the remaining limb D of the T tube attach a rubber tube, clip, and short glass tube. Have a 500 c.c. flask with one-holed cork ready to attach to this

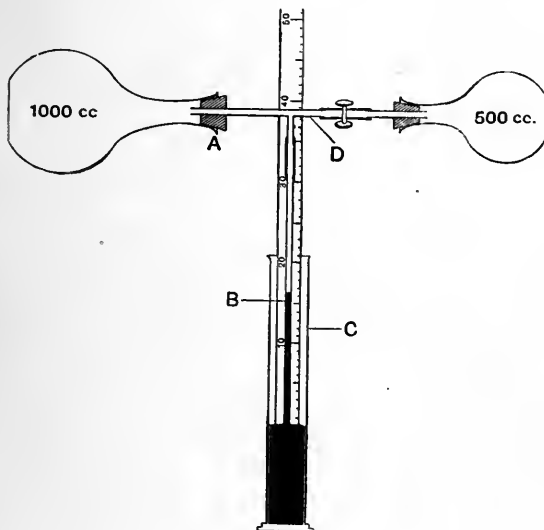


FIG. 127.

tube. Open the clip. Suck at the short glass tube attached to D. The mercury rises in the pressure-gauge. Raise the mercury in this way as high as you can. Close the clip. Read the level of the mercury. Fill the 500 c.c. flask with coal gas. Attach it to the short glass tube. Open the clip. Some coal gas enters the 1,000 c.c. flask. The mercury falls. Read the pressure-gauge, and also the barometer. The volume of each flask up to the cork must also be known or found.

## EXAMPLE.

Barometer ... ..	726 mm.
Pressure-gauge (1) ... ..	127 mm.
(2) ... ..	81 mm.
Volume of large flask ... ..	1000 c.c.
Volume of small flask ... ..	500 c.c.

Hence,

Small flask contains 500 c.c. gas at 726 mm.

Large flask contains 1,000 c.c. air at (726 - 127) or 599 mm.

If 1,000 c.c. air expanded to fill 1,500 c.c., its pressure would become

$$599 \times \frac{1000}{1500} \text{ or } 399.6 \text{ mm.}$$

If 500 c.c. gas expanded to fill 1500 c.c., its pressure would become

$$726 \times \frac{500}{1500} \text{ or } 242 \text{ mm.}$$

Now,

observed pressure of mixture is 726 - 81 or 645 mm.

What relation is there between the numbers 399.6, 242, and 645?

We notice that by *addition* 399.6 + 242 are equal to 641.6, or so nearly 645 that the difference is probably only due to errors in the experiment.

Hence, the observed pressure of the mixture is shown by this experiment to be equal to the sum of the calculated partial pressures of the air and coal gas.

It seems then that the pressure exerted by either the air or coal gas is quite unaffected by and independent of the presence of the other gas.

It would of course be necessary to make many more experiments to assure ourselves that what we have found to be true of this mixture of gases is true of all others. We may, however now accept the conclusion at which Dalton arrived.

**Dalton's Law of Mixed Gases.**—*When several gases or vapours are present in the same space, the pressure exerted by the mixture is the sum of the partial pressures exerted by the individual gases or vapours.*

We may state this in mathematical form as follows :—

Let  $P$  = pressure exerted by the mixture.

$P_1$  = partial pressure exerted by one gas or vapour.

$P_2$  =           "           "           a second           "

$P_3$  =           "           "           a third           "

&c.                                   &c.

Then,

$$P = P_1 + P_2 + P_3 + \dots$$

**ILLUSTRATION.**—The pressure of the atmosphere must be thought of as equal to the sum of the partial pressures of the oxygen, nitrogen, water vapour, and any other gases which it may contain.

**EXAMPLE.**—Find the partial pressures in an atmosphere consisting of 4 volumes nitrogen and 1 volume oxygen when the total pressure is 750 mm.

If 4 volumes nitrogen expanded to 5 volumes, its pressure would be  $\frac{4}{5}$  of 750 or 600 mm.

If 1 volume oxygen expanded to 5 volumes, its pressure would be  $\frac{1}{5}$  of 750 or 150 mm.

**Correction of the Volume of a Gas for Temperature, Pressure, and Water Vapour.**—The volumes of gases collected over water must be corrected not only for temperature and pressure, but also for water vapour. They become saturated with water vapour, and therefore more bulky ; but the correction is best made by thinking of the changed pressure. The pressures of saturated water vapour at different temperatures may be found as described on page 85. The results are :—

Temperature ... ..	0°	5°	10°	15°	20° C.
Pressure of water vapour	4 +	6 +	9 +	12 +	17 + mm.

**EXAMPLE.**—Suppose 100 c.c. of dry air at 0° and 760 mm. were enclosed over water, what would the volume become at 750 mm. and 15° C. ?

(i) *Correction of Pressure for Water Vapour.*

The pressure of the atmosphere would be balanced by the sum of the partial pressures of the enclosed air and the water

vapour mixed with it. Then the pressure of the enclosed *air* would be less than that of the atmosphere. From the tabulated results, at  $15^{\circ}\text{C}$ . pressure of water vapour is 12 mm. Hence, pressure of enclosed air would be  $750 - 12$  or 738 mm.

(ii) *Correction of Volume for Pressure* . . .  $\times \frac{760}{738}$

(iii) *Correction of Volume for Temperature.*

$15^{\circ}\text{C} = 288^{\circ}\text{A}$ . Correction . . .  $\times \frac{288}{273}$

Hence, *corrected Volume* is  $100 \times \frac{288}{273} \times \frac{760}{738}$   
 $= 108.6 \text{ c.c.}$

### A Self-Acting Instrument for supplying at once a Correction for Temperature and Pressure and Water

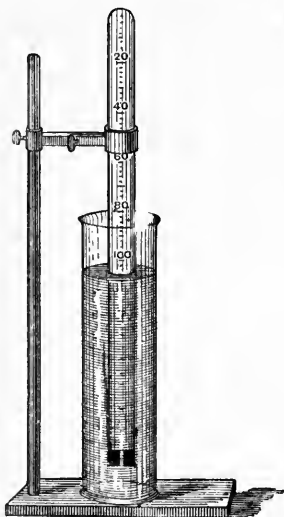


FIG. 128.

**Vapour.**—If 100 c.c. of dry air at N.T.P. had been introduced into a tube over water, it would as just shown occupy 108.6 c.c. at  $15^{\circ}\text{C}$ . and 750 mm. Suppose these are the conditions of the laboratory. If we put 108.6 c.c. of wet air into a tube over water then this contains the *mass* of air which will return to 100 c.c. when dry at N.T.P.

Obtain or make a tube enclosed at one end, and graduated in c.c. from 100 to 120 c.c. (Tubes may be bought ready-made and graduated from 0 to 200 c.c.) Fill the tube half full with water. Fit it with a one-holed cork. Invert it in a tall jar of water. Immerse it until the liquid levels are equal within and without (Fig. 128). Note the volume of the contained

air. Admit air bubble by bubble until just 108.6 c.c. are shut up. The instrument is complete for reference.

Suppose that on a later day the temperature and pressure

change. The volume of air we enclosed in the tube is now found to be for example 109 c.c. Then any volumes of moist gases measured that day should be corrected by  $\times \frac{100}{109}$ .

### PROBLEMS

1. Find whether there is any change of volume when oxygen and nitric oxide are mixed.
2. Measure the change of volume on mixing 100 c.c. water with 100 c.c. spirits.
3. Read the barometer and thermometer. Set up a gas-correcting instrument as described above.

### EXAMPLES XVI

#### Dalton's Law of Mixed Gases

1. Two flasks containing respectively 1000 c.c. of air at 750 mm. and 500 c.c. of coal gas at 600 mm. are connected. Find the pressure of the mixture.
  2. Into a nearly vacuum flask of 1000 c.c. capacity containing air at 10 mm., are admitted 10 c.c. of oxygen at 750 mm., and 20 c.c. of hydrogen at 720 mm. Find the total pressure.
  3. What is the partial pressure of the nitrogen in the atmosphere when the barometer reads 750 mm. ? (Nitrogen = 79 per cent. of air by volume.)
  4. Find the sum of the partial pressures of 5 c.c. oxygen at 720 mm. and 10 c.c. hydrogen at 600 mm., when allowed to enter a flask of 100 c.c. capacity.
  5. Find the volume of a mixture of 4 litres nitrogen at 760 mm. and 1 litre oxygen at 720 mm., when mixed at a pressure of 750 mm.
  6. Find the volume *dry* at N.T.P. of 1000 c.c. of hydrogen saturated with moisture at 760 mm. and 15° C.
  7. Find the volume at N.T.P. of 100 c.c. of dry air collected over mercury at 750 mm. and 10° C., the mercury standing 20 mm. higher inside the tube than without.
  8. Find the volume when *dry* at N.T.P. of 100 c.c. of air measured at 760 mm. and 0° C., which is standing over water 2.6 c.m. lower inside the tube than without. Relative density of mercury 13.6.
- See p. 263 for Pressure of Aqueous Vapour.

## EXAMPLES XVII

## The Densities of Gases

Assume—1 litre of hydrogen weighs '09 gram at N.T.P. ; also densities relative to hydrogen :—

Oxygen ... ..	16	Steam ... ..	9
Nitrogen ... ..	14	Carbonic acid ...	22.

1. Find the weight of a litre of oxygen at N.T.P.
2. What volume will 1 gram of carbonic acid gas fill?
3. Find the volume occupied by 1 gram of steam at  $91^{\circ}\text{C}$ .
4. Find the weight of a litre of steam at  $104^{\circ}\text{C}$
5. Find the volume occupied by 169 grams of carbonic acid at  $51^{\circ}\text{C}$ . and 741 mm.
6. Find the weight of 19 c.c. oxygen at  $13^{\circ}\text{C}$ . and 770 mm.
7. Find the volume at N.T.P. of a mixture of '23 gram oxygen and '77 gram nitrogen.
8. Find the weight of a mixture of 210 c.c. oxygen and 790 c.c. nitrogen.
9. Find the relative density (compared to hydrogen) of a mixture of 20'61 litres oxygen, 77'95 litres nitrogen, '04 litre carbonic acid and 1'40 litres steam.
10. Find the number of grams of oxygen present with each gram of hydrogen in a mixture which contains equal volumes of each.
11. Find the partial pressure in a mixture of equal weights of oxygen and hydrogen. Pressure of mixture = 765 mm.
12. Find the pressure of steam due to admitting 1-gram of water to an empty flask of 1 litre capacity at  $132^{\circ}\text{C}$ .

## CHAPTER XXX

### A RESEARCH ON CHALK

WE shall now return to Black's research<sup>1</sup> on chalk, and endeavour to follow those lines of inquiry which Black himself pursued. We may hope to arrive at important results by making in this way a thorough examination of one substance, particularly if we pay close attention to the quantities of the materials with which we are dealing.

Chalk is a substance which causes the liberation, as we have already found, of carbonic acid gas when hydrochloric acid is added to it, and yields lime when it is heated in a kiln. This is also true of calcspar and marble. Pending further evidence we shall use the term chalk as including all of these.

### THE ANALYSIS OF CHALK.

(a) THE ACTION OF ACIDS UPON CHALK.—EXPT. I. To find the Weight of Carbonic Acid expelled from 100 Grams of Chalk in contact with Hydrochloric Acid.—We might attempt to collect and weigh the gas itself, but it is easier to find the *loss in weight* due to escaping gas. Choose a small lipped beaker and a small flask, small enough to stand side by side upon the balance pan.

<sup>1</sup> A reprint of Black's admirable paper (1755) may be obtained in No. 1 of the "Alembic Club Reprints," W. F. Clay, Edinburgh, 1s. 6d. net. Advanced students and teachers of Chemistry are strongly recommended to obtain it. A few years later Lavoisier repeated and extended Black's experiments. He gave in 1773 a very clear account of them and of the history of the discovery of carbonic acid gas. A translation of this by Henry under the title of *Essays, Physical and Chemical*, is very readable and well deserves perusal.

Into the beaker pour about 15 c.c. of strong hydrochloric acid. It fumes and may lose weight by fuming. Add about 45 c.c. of water; the fuming stops. Place a short glass tube, which has been narrowed at one end, in the beaker. Weigh the half-full beaker and the empty flask together. Place about two grams of *dry* chalk<sup>1</sup> in the flask. Weigh the flask and beaker again (Fig. 129). Transfer carefully a little acid by means of the



FIG. 129.

dropping tube from the beaker into the flask. There is an effervescence of carbonic acid gas. We wish to liberate all the carbonic acid; so go on adding acid until effervescence ceases. By means of the tube suck out carbonic acid from the flask. Weigh the beaker and flask again together with the tube. There is a decrease in weight. Calculate the loss

in weight there would be in the case of 100 grams of chalk. What causes of error are there possibly in this experiment?

In successive experiments performed by pairs of boys the following numbers were obtained for the percentage of carbonic acid expelled from chalk and hydrochloric acid:—45, 42·4, 45·4, 43, 45, 43. *Mean*: = 43·9. Do your experiments lead to the same result?

Careful experiments show that whatever the quantity of the chalk that is taken, there is always a loss of about 44 per cent. This shows that the quantity of carbonic acid evolved depends upon the quantity of chalk taken, and not upon the quantity of acid added. Is it therefore certain that the carbonic acid comes from the chalk and not from the acid (*vide* p. 158)?

(b) THE EFFECT OF HEAT UPON CHALK.—As we have already learnt, it has long been known that when chalk is exposed to a violent heat it is converted into a substance called quicklime. The alchemists and early chemists regarded this as a peculiar acrid and caustic earth, formed by the union of the chalk with a fiery substance.

In limestone and chalk districts you may have seen kilns in which quicklime is obtained by heating the stone with a little coal. The change can, however, be brought about by heat alone.

<sup>1</sup> Use powdered calcspar, or precipitated chalk, B.P., dried at 100°.



**EXPT. 2. To find whether Chalk is changed by Heat Alone.**—(a) Damp a slip of red litmus paper. Let a little chalk dust fall upon it. The colour is unchanged.

(b) Heat strongly a little chalk dust on platinum foil with the Bunsen flame. Then drop it on the damp test-paper. The colour goes blue where the heated chalk has touched it. Then the chalk dust has, in part at any rate, been converted into quicklime.

It had apparently occurred to no one before the time of Black to really examine what changes were effected; no one had thought of *weighing* the chalk before and after it was heated.

**EXPT. 3. To find the Weight of Quicklime obtained from 100 Grams of Chalk by Heat.**—Weigh a small porcelain crucible.<sup>1</sup> Weigh into it from one to two grams of pure chalk. Place it in a muffle furnace,<sup>2</sup> and heat strongly for half an hour. Remove the crucible from the furnace. Allow it to cool. Weigh. Reheat for ten minutes and re-weigh, and so on till there is no further change in weight. Here is the result of an actual experiment :—

Weight of crucible	...	...	...	...	...	11.705	grams.
Weight of crucible and chalk	...	...	...	...	...	12.841	,,
Weight of chalk	...	...	...	...	...	1.136	,,
Weight of crucible and lime	...	...	...	...	...	12.338	,,
Weight of lime	...	...	...	...	...	0.633	,,
Loss of weight of chalk	...	=	1.136	-	0.633	or	0.503

$$\text{Percentage loss of weight of the chalk} = \frac{0.503}{1.136} \times 100 \text{ or } 44.3$$

Pure chalk is always found to lose about 44 per cent. in weight, when ignited (*i.e.* strongly heated). Thus, instead of gaining anything from the fire, the chalk loses the *same* amount as when treated with an acid. It appears, therefore, from these experiments that the carbonic acid which is expelled from chalk and an acid is also driven off by heat, when chalk is converted

<sup>1</sup> A platinum crucible and a blow-pipe flame may be used with advantage if available.

<sup>2</sup> Fletcher's Reverberatory Gas Furnace, 23H is suitable for class work. A number of crucibles can be heated in it at the same time.

into quicklime. It is therefore clear that in the former case the carbonic acid comes from the chalk.

If this is so, when acid is added to the quicklime there will be no effervescence.

**EXPT. 4. To try whether Quicklime contains Carbonic Acid.**—Drop hydrochloric acid on to a fragment of freshly burnt quicklime in a test-tube. The drops hiss as they touch the lime. Cover the lime with more acid. The liquid soon becomes hot and bubbles. Are the bubbles steam or gas? Cool the tube by plunging it into cold water. The bubbling stops. Try the experiment again using weak instead of strong acid. The mixture gets warm—not so hot as before, and there are no bubbles. The bubbles seem to have been steam, not gas; the mixture, then, did not effervesce, but boiled. Notice that the lime is gradually dissolved by the acid.

Quicklime, then, is chalk which has been deprived of its carbonic acid.

#### CONCLUSION FROM THE ANALYSIS OF CHALK—

Now, 100 grams chalk lose on solution in acid 44 grams in weight,  
and at the same time carbonic acid is evolved.

100 grams chalk lose on ignition 44 grams in weight.

100 grams chalk yield on ignition 56 grams quicklime, which  
contains no carbonic acid.

But  $44 + 56 = 100$ .

Therefore, there seems good reason to believe that

*chalk = lime + carbonic acid gas,*<sup>1</sup>

nothing more and nothing less.

**The Properties of Quicklime and of Slaked Lime.**—**Preparation of Pure Quicklime.**—We shall find presently that quicklime deteriorates when exposed to the air. On this account quicklime, when not fresh, should be reburnt before experimenting with it.

Put some lumps of quicklime into the middle of a hot coal fire. Leave them there red-hot for several hours. Then remove them. To test the purity of the lime add water and hydrochloric acid. There should be no effervescence.

<sup>1</sup> The sign = means 'yields' or 'gives.' It has not the same meaning as in algebra, and chemical equations (unlike algebraic equations) read only from left to right.

Place the rest in tin boxes with closely fitting lids, for future use. Seal with paraffin. Even then the quicklime should be reburnt before any exact experiment. The lightest pieces are generally the best. Heavy and stony pieces should be rejected.

**EXPT. 5. To Slake Quicklime.**—It has long been known that water has an action upon quicklime. Place a lump of quicklime upon a plate. Pour water over it. Observe that it presently crumbles up: much heat is evolved, for steam is produced. The quicklime is said to be *slaked*, or converted into slaked lime. What has occurred? Suppose we ask the balance.

**EXPT. 6. Slaked Lime is Quicklime and Water.**—Take a small crucible. Weigh it. Introduce about 2 grams of good fresh quicklime. Weigh. Carefully moisten it with about 1·5 c.c. of distilled water. Place the crucible in a steam-jacketed oven, or in an oven kept at 100° C., in order to evaporate excess of water. Weigh. Reheat at 100°, and reweigh, till no further decrease in weight occurs. In an actual experiment:—

Weight of crucible	... ..	13·331	gram.
Weight of crucible and quicklime	... ..	13·580	gram.
Weight of quicklime	... ..	0·249	gram.
Weight of crucible and slaked lime	... ..	13·663	gram.
Weight of slaked lime	... ..	0·332	gram.

Gain in weight of slaked lime = 0·332 - 0·249 or 0·083 gram.

$$\begin{array}{ccccccc} \text{,,} & \text{,,} & \text{,,} & \text{,,} & = & \frac{0\cdot083}{0\cdot249} & \times 100 \text{ or } 33\cdot3 \text{ per cent.} \end{array}$$

Then, when quicklime is slaked with water, it *gains* in weight.

When good quicklime is carefully slaked, there is always found a gain in weight of about 32 per cent. This must be due to water uniting with the quicklime.

Hence,

$$\text{Slaked lime} = \text{quicklime} + \text{water.}$$

We found some time ago that slaked lime is slightly soluble in water. But we have not obtained as yet an answer to the following question.

**EXPT. 7. Can Slaked Lime be entirely dissolved in a Large Quantity of Water?**—Why does not the white mud of slaked lime entirely dissolve in water? Is the mud which settles something different from the lime which dissolves—*i.e.* is the lime impure? Does it need heat, or shaking, or longer time? Try the experiment again, giving the lime every chance. Choose a good piece of lime—one which feels light—or one of the lime cylinders used for the lime-light. Burn it in a coal fire, and when still fresh take not more than 0.5 gram. Have at least 500 c.c. of freshly boiled distilled water ready in a flask. Slake the lime with a few c.c. of boiling hot distilled water. Throw the mixture at once into the flask of hot water. Shake.

Black tried this experiment and found that the whole of the lime dissolved with the exception of a very small amount of sediment. This, as Black observed, was probably due to the accidental admixture of a little foreign matter.

**Is Quicklime an Element?**—Is quicklime an elementary substance, *i.e.*, an element? So far as our present study of the subject goes, and so far as Black knew, it is an element, for we do not know how to break it up into different substances. But observe that this does not prove that it is really an element. Indeed, quicklime resembles magnesia very closely in appearance and properties. Both are infusible, white, earth-like substances; and magnesia can be obtained from another white substance, “magnesia alba,” in just the same way as quicklime from chalk. Up to the time of Black, these substances were frequently confused with one another. We know now that magnesia is really an oxide of the metal magnesium. It is possible, then, that further study or research may prove quicklime to be an *oxide* of a metal.

**THE SYNTHESIS OR PREPARATION OF CHALK FROM LIME.**—We have shown by analysis that chalk is a compound of lime and carbonic acid gas. Can we not then perhaps produce chalk by uniting carbonic acid with lime?

**EXPT. 8. To find whether Chalk can be formed from Lime and Carbonic Acid.**—Powder some quicklime. To part of the powder add water, and then hydrochloric acid. If it does not effervesce, it is free from chalk, and will do for this

experiment. Place about 1 gram of the quicklime in a test-tube. Half fill the tube with water. Bubble carbonic acid gas up through the milky liquid for a few minutes. Pour the whole of the contents of the tube upon a filter. Wash two or three times with water. Drop 2 or 3 drops of hydrochloric acid on to the white substance on the paper. It effervesces violently. Then chalk has probably been formed.

How could it be proved beyond any doubt that it is chalk, and not anything else containing carbonic acid, which has been formed? Would it not be well to find whether the product contains the same percentage of carbonic acid that we know to be present in chalk? Proceed thus. Pass carbonic acid gas for half an hour through water in which lime is suspended. Boil for a few minutes. Filter. Dry the filter-paper with the precipitate in a steam oven at 100 degrees. Heat some of the dry substance in a crucible as in Experiment 3, and find the percentage loss of weight. It is nearly 44 per cent.

We have, then, succeeded in forming chalk by putting together lime and carbonic acid gas. Expressed briefly, in addition to the previous conclusion that

$$\text{Chalk} = \text{lime} + \text{carbonic acid},$$

we have now shown that

$$\text{Carbonic acid gas} + \text{lime} = \text{chalk}.$$

This is another instance of what is called a *synthesis*, or placing together. It affords most important additional evidence of the nature of chalk, and confirms the conclusion we had arrived at from the *analysis* of chalk.

The last experiment renders it extremely probable that it is chalk which is produced when carbonic acid is passed through clear lime water. (See Problem 1, p. 156.)

**EXPT. 9. The Lime Water Test for Carbonic Acid.**—Pour a little lime water into a jar containing carbonic acid gas. Shake. It becomes milky. A slight precipitate forms. Show that it is chalk.

Clear lime water, as we found previously, may be used as a delicate test for carbonic acid. We now understand that the lime water is rendered milky because chalk is at once formed.

which is insoluble in water. One drop of lime water is often enough. For example :—

**EXPT. 10. To show that Carbonic Acid is given off when Chalk is heated.**—Place a few grams of chalk in a hard glass test-tube. Heat the tube in the foot-blowpipe flame. Dip a glass rod into lime water. One drop of lime water hangs at the end of the rod. Carefully lower the drop into the test-tube, so that it does not touch the sides. Withdraw the rod. Examine the drop against a dark background. It has become cloudy.

Can chalk be prepared from lime in any other way? Now, washing soda (milk alkali) contains carbonic acid (*see* Problem 1, p. 159). What then may the effect be of adding quicklime to a solution of washing soda? May not the lime recover carbonic acid from the soda? This is a question Black asked himself. We may answer it by experiment, as he did.

**EXPT. 11. To find whether Chalk can be reformed from Lime by means of Washing Soda.**—Weigh out about 7 grams of washing soda. Dissolve the soda in about 60 c.c. water. Filter the solution into a flask. Powder finely about 1 gram of pure quicklime. Throw the quicklime into the solution of soda. Warm and shake well. Pour the product on to a filter-paper. Wash the precipitate on the paper repeatedly with hot water, so as to free it completely from the soda. Pour a few drops of acid on to the paper. There is a violent effervescence. The precipitate contains, therefore, ordinary chalk. Then the carbonic acid which the lime has gained must have been furnished by the washing soda.<sup>1</sup>

**EXPT. 12. To find the Weight of Chalk obtained from 56 Grams of Quicklime dissolved in Acid.**—Weigh out accurately into a flask 1.5 to 2 grams of pure quicklime. (From what quantity of chalk could it be obtained?) Add water and hydrochloric acid. If there is no effervescence, it proves that the lime is free from gas. If there is any effervescence, then any gas contained as an impurity in the lime is got rid of. The lime will dissolve on warming. If it does not

<sup>1</sup> Advanced students may determine the quantity of chalk produced by this means from a weighed quantity of quicklime. The precipitate should be dried in an oven at 100°.

all disappear, add a little more acid. The solution now contains excess of acid. Add to it a clear solution of washing soda. There is an effervescence. (Why?) Then a precipitate forms. But on shaking, this precipitate disappears with effervescence. (Why?) After adding still more soda the effervescence stops (why?), and the precipitate remains undissolved. Go on adding more soda until no more precipitate forms. To know when there is enough soda, filter a little of the liquor containing the precipitate through a counterpoised filter-paper. Add soda to the filtrate. If there is no precipitate, then the soda is already in excess.

The precipitate must now be obtained clean from the soda, hydrochloric acid, and other substances which may be present. Heat the liquid until it boils. The precipitate clots and settles. Pour the liquid on to the filter used before. Wash the precipitate again with hot water, and pour the washings on to the filter. Pour the precipitate on to the filter. Go on washing it with hot water, until it is quite free from soda or other dissolved substances. Put one drop of the filtrate as it drips from the funnel on to platinum foil and evaporate. If it leaves a white mark the precipitate is not yet clean, and must be washed with more water.

Dry completely the filter-paper with the chalk in the funnel at  $100^{\circ}$  in an oven. Carefully place the paper with the chalk on the pan of the balance, place the counterpoise of the filter on the other pan and weigh again. Find the weight of the chalk alone. Hence, calculate the weight of chalk obtained from 56 grams of quicklime. Compare your results with those obtained in Experiment 3.

**EXPT. 13. An Attempt to recover Lime from its Solution in Acid.**—Lime dissolves in acid. Chalk when dissolved in acid loses carbonic acid gas. Can we recover the lime by evaporating the acid?

Leave some chalk soaking in hydrochloric acid so that as much chalk may dissolve as possible. Filter the liquid. Evaporate in a porcelain basin to dryness.

Is the solid white substance lime? Try whether a portion forms lime-water with water. No! It dissolves much more readily. Is the solution alkaline? Notice that the solid quickly

becomes sticky and moist if left exposed to the air. Evidently it has the power of attracting moisture from the air.

This substance is called "*calcium chloride*." Its most remarkable property is the readiness with which it picks up moisture. Hence, it is a useful drying agent.

**The Volume of Gas expelled from Chalk.**—We have found that a definite quantity of chalk loses, on ignition or on solution in acids, a fixed *weight* of carbonic acid gas. Hence, a fixed *volume* of the gas should be obtained from a definite quantity of chalk.

EXPT. 14.<sup>1</sup> To find the Volume in c.c. of Carbonic Acid Gas expelled from One Gram Chalk by Hydrochloric Acid.—Weigh out carefully about 0.25 gram of chalk upon a piece of paper fold this and introduce it into a 2 oz. flask *a*. Add about 5 c.c. of water. Carefully introduce a small glass tube containing about 5 c.c. of dilute hydrochloric acid suspended by a thread; close the flask with a tightly-fitting india-rubber stopper, through which is passed a glass tube connected with two

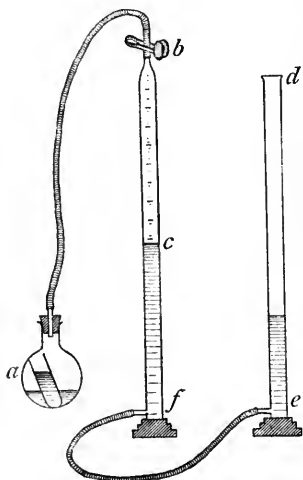


FIG. 130.

feet of thin rubber tubing. Then attach the free end (Fig. 130) of the tubing to a pair of Hempel's burettes filled with water. Place the burette *de* at a lower level than the other, and then open the clamp *b*. Tilt the flask so that the acid is brought into contact with the chalk. Carbonic acid is immediately evolved and forces the water down the tube *bc*. If the flask is gently shaken the chalk will entirely dissolve.

Finally raise the burette *de* until the water stands at the same

<sup>1</sup> Lecture Table Experiment.



level in both columns. Read off the volume of air or gas in *bc*, and calculate the volume of carbonic acid which 100 grams of chalk would evolve. In an experiment :—

0.25 gram chalk evolved 55 c. c. carbonic acid.

Hence,

1           ,,           ,,           220 c.c.           ,,

The experiment should be repeated several times. It is found that at 15° C. about 230 c.c. of carbonic acid are always obtained from one gram of chalk.

### PROBLEM

Fit up the apparatus shown in Fig. 130. Use an 8 oz. flask. Find how to use it so as to measure the volume of carbonic acid gas expelled from chalk when dissolved in dilute hydrochloric acid. Take about 1 gram of chalk. Is there any source of error?

**The Atmosphere contains Carbonic Acid Gas.**—You must have noticed that when strong lime water is exposed to the open air a pellicle or crust forms upon its surface. Try also *sucking* a current of air through lime water placed in a wash-bottle. What can the thin white pellicle be which is formed also in this case? If collected and examined it is found to be chalk. How can the chalk have been formed? Only through the union of the lime with carbonic acid gas. The carbonic acid can only have come from the atmosphere. We have then arrived at the important conclusion that the atmosphere contains floating about in it particles of carbonic acid. As Black said in 1755 :—

“If this fluid be exposed to the open air, the particles of quicklime which are nearest the surface, gradually attract the particles of Fixed Air which float in the atmosphere. But at the same time that a particle of lime is thus saturated with air, it is also restored to its native state of mildness and of insolubility, and as the whole of this change must happen at the surface, the whole of the lime is successively collected there, under its original form of an insipid calcareous earth, called the cream or crust of lime water.”

Carbonic acid gas can always be found in the atmosphere, though usually only in small quantities. There are however

some places where it is found in large amounts ; in particular the Poison Valley in Java, an old volcanic crater, and the Dog Grotto near Naples, where it seems to arise from rents in the floor of the cave. It is also found abundantly in mines and old wells.

### The Composition of Mild Alkali or Washing Soda.

We have found that chalk is precipitated when lime is mixed with a solution of washing soda, or of "mild alkali," as Black called it. Now washing soda contains carbonic acid. When chalk is produced in this way, carbonic acid is transferred from the washing soda to the lime. What is left of the washing soda? It was stated on p. 139, that the substance called "caustic soda" could be prepared by these means. But until we have ourselves prepared it, suppose we call this other constituent of the washing soda X. Then :—

$$\text{Washing soda} = \text{X} + \text{carbonic acid.}$$

Then we can also express the change which occurs when lime is mixed with a solution of washing soda thus :—

$$\text{Lime} + \text{washing soda} = \text{chalk} + \text{X.}$$

*i.e.*

$$\text{Lime} + (\text{X} + \text{carbonic acid}) = (\text{lime} + \text{carbonic acid}) + \text{X.}$$

This is a chemical equation. What we require to do is to obtain the substance X.

**EXPT. 15. To obtain the Other Constituent of Washing Soda.**—Weigh out about 10 grams of quicklime. Slake it with boiling water. Then stir it up with more water so as to form a milky liquid of about 50 c.c. volume. Dissolve about 10 grams washing soda in 50 c.c. hot water. Add the solution to the milky liquid, stirring at the same time. Boil for ten minutes, and replace from time to time the water which evaporates. The white mud may still be lime for all we have yet learned, or it may have been changed to chalk, or it may be partly lime and partly chalk. We must keep in mind that besides the visible mud there may be other substances present in solution. Washing soda may be in solution ; lime may be in solution. The substance X may be either in the solution or in the mud. Is it likely that both washing soda and lime are

in solution together? No, since their solutions yield a precipitate when mixed (*see* Experiment 11).

Let us find whether the solution contains any washing soda. Remove the flame. The mud begins to settle. Filter a few drops of the liquid. Add a drop of hydrochloric acid to the filtrate. It does not effervesce. Then there is no washing soda in the solution. (If there is any effervescence the liquid must be boiled longer.) The precipitate is so bulky, that it will hardly separate from the liquid. Pour both into a tall jar. Add an equal volume of water. This dilutes the solution. But on settling the precipitate occupies relatively less room, and allows the greater part of the lye to be withdrawn as a clear liquid.

**EXPT. 16. The Lye is Caustic.**—Place a few drops of the lye in a small smooth-surfaced porcelain crucible. Evaporate by heating the crucible on a sand-bath, surrounding it well with the sand. Cool. It solidifies as a white mass. Re-dissolve in water and wash out the basin. Observe that its inner surface is corroded. As Black said, "The acrimony of the lye is so great that having once evaporated a part in a bowl of English earthen or stone ware . . . it corroded and dissolved a part of the bowl, and left the inside of it pitted with small holes." On this account it is customary to evaporate the lye, when desired, in a silver or iron<sup>1</sup> dish, which are found to remain unattacked.

If a little of the lye is placed upon the skin it will quickly raise a blister. On account of these properties it was called by Black and others "*caustic alkali*," in distinction from "mild alkali" (washing soda). It is now called "*caustic soda*."

Have we proved that X is caustic soda alone?

When mild alkali (washing soda) is added in sufficient amount to lime dissolved in hydrochloric acid chalk is precipitated, as Experiment 11 showed. What will be the result of using caustic alkali in place of the mild alkali? Chalk cannot be formed. Why not? But the caustic alkali will neutralise the acid in which the lime is dissolved. Then perhaps the lime will be precipitated? Try.

**EXPT. 17. To find whether Lime is precipitated when Caustic Alkali is added to Lime dissolved in Hydrochloric Acid.**—Place about 1 gram of pure quicklime

<sup>1</sup> Platinum or aluminium dishes must not be used.

in a flask. Add water and hydrochloric acid. Warm till the lime is dissolved. Add a clear solution of caustic soda. A white precipitate is formed. Add the caustic soda until no further precipitate is formed. Heat until the liquid boils. Wash by decantation and then pour the contents on to a filter-paper, as in Experiment 12. Wash thoroughly with hot water, to free it from caustic soda. Dry it at  $100^{\circ}$  in an oven. Examine the product. Lavoisier performed this experiment, and wrote :—

“This earth was a true lime ; it dissolved in water, in the same proportion as lime. The lime water prepared with it threw up a cream of chalk on its surface ; it made scarcely any effervescence with acids ; it communicated causticity to alkalies ; in short, no difference could be perceived between this and real lime prepared by calcination.”

The results of our study show that quicklime and caustic soda are less composite substances than chalk and washing soda. Yet the very reverse was supposed to be the case before the work of Black which we have been closely following. For the acrimony of the caustic substances was supposed to be due to the presence of a fiery substance, absent from the mild substances. Black sums up the matter thus : “If quicklime be mixed with a dissolved mild alkali, it shows an attraction for Fixed Air superior to that of the alkali. It robs this salt of the air and thereby becomes mild itself ; while the alkali is thereby rendered more corrosive.”

### PROBLEMS

1. Find whether chalk, plaster of Paris, bicarbonate of soda are soluble in water or not.
2. Are the substances X, Y, Z, chalk or lime, or both or neither ?
3. Invent a substitute for the Hempel burette, using the ordinary apparatus of the laboratory.
4. Find whether there is any loss in weight when quicklime is dissolved in hydrochloric acid.
5. Can lime be made by heating chalk in an ordinary coal fire ?
6. What weight of slaked lime can be dissolved by a litre of water ?
7. Find the percentage loss of weight when dry slaked lime is strongly heated.

8. Find whether the given sample of "quicklime" contains (a) any slaked lime, (b) any quicklime.

9. Find the effect of passing a long continued current of carbonic acid gas through lime water.

10. Find whether slaked lime is changed to chalk by moistening it with a solution of washing soda.

11. What happens when a solution of washing soda is added to lime water?

12. Find the effect of adding lime water to water saturated with carbonic acid gas and reddened by litmus.

13. Find whether the given sample of caustic soda contains (a) any Fixed Air, (b) any lime.

14. Weigh out accurately about 3 grams pure quicklime. Prepare chalk from it following the method described in Experiment 12. Estimate the weight of chalk produced.

15. You are provided with some caustic alkali (caustic soda). Prepare some mild alkali (washing soda) from it.

16. Find in the case of (a) crystals of washing soda, (b) carbonate of magnesia, (c) carbonate of copper :—

(i) The percentage loss in weight when treated with hydrochloric acid.

(ii) The percentage loss in weight when strongly heated.

(iii) The effect of hydrochloric acid upon the substances obtained by ignition as in (ii).

(iv) The volume of carbonic acid expelled from 1 gram on treatment with hydrochloric acid.

17. Find the percentage weight of carbonic acid which can be obtained from Dover chalk, mountain limestone, coral, oyster-shells, egg-shells.

18. Prepare a sample of carbonate of magnesia from magnesia.

19. Prepare a sample of carbonate of copper from oxide of copper.

20. Dissolve in water the white powder obtained on heating washing soda, evaporate the solution, and identify the crystals which are obtained.

21. Dissolve some iron nails in hydrochloric acid. Filter from any undissolved matter. Find the effect of adding to the solution (a) a solution of washing soda, (b) a solution of caustic soda. Collect, dry, and examine the products.

## CHAPTER XXXI

### THE COMPOSITION OF WATER BY VOLUME AND BY WEIGHT

WE have already established beyond doubt the fact that water is a compound of hydrogen and oxygen, and of nothing else. Moreover, we already have some evidence as to the relative volumes and the relative weights of hydrogen and oxygen which combine together when water is formed, or are liberated when water is decomposed. This question of the quantitative composition of water must now be considered more fully.

#### A. THE COMPOSITION OF WATER BY VOLUME.

There are two distinct lines of inquiry to be followed, just as when we were investigating the composition of air, and that of carbonic acid gas.

1. ANALYSIS.—We have already learnt that when water is decomposed by the electric current, two volumes of hydrogen are liberated for every one volume of oxygen. So far then as the analysis of water can afford evidence, the answer is definite that water consists of two volumes of hydrogen combined with one volume of oxygen.

2. SYNTHESIS.—*Priestley* found that when hydrogen was exploded with oxygen in a bottle, the loudest explosion was obtained when about two-thirds of the bottle was filled with hydrogen and the rest with oxygen.

*Cavendish* showed in 1781 that 1000 volumes of air always contain nearly 210 volumes of oxygen. He abstracted the oxygen with nitric oxide, just as Dalton had done. He arrived at the conclusion that common air contains 20.83 per cent. of oxygen by volume. In the same year he went on to explode common air with different proportions of hydrogen (*see* p. 219). He found that it was necessary to add 423 volumes of hydrogen to 1000 volumes of common air in order to burn the whole of the oxygen in the air. Almost all the hydrogen too was consumed. His experiment, then, proved that the water formed consisted of

$$\begin{cases} 210 \text{ volumes of oxygen.} \\ 420 \text{ volumes of hydrogen.} \end{cases}$$

In other words, one volume of oxygen and two volumes of hydrogen united together forming water. Although the experiments of Cavendish justified this conclusion, he did not state it in this form.

Further experiments carried out by the French chemist Gay Lussac and by Von Humboldt were necessary to lead chemists to regard the composition of water by volume established.

**EXPT. <sup>1</sup> I. Two Volumes of Hydrogen are combined with One Volume of Oxygen in forming Water.**

(1) *Obtain a U-shaped "eudiometer."*

This is a U-shaped tube, one limb of which is closed by a glass tap *a* (Fig. 131). Two platinum wires are fused into the glass near the top of the limb *bc*, and the limb may be graduated or plain. A tap *e* is fixed to the tube near the bottom of the other limb. Close the tap *e* and open the tap *a*. Pour mercury cautiously down the limb *de*. When the mercury has risen to *a* and expelled all the air, close the tap *a*. Open the tap *e* and run out the mercury slowly. The level of the mer-

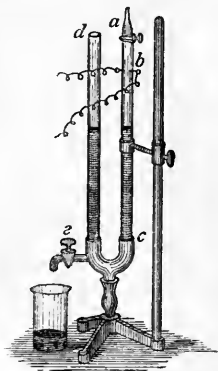


FIG. 131.

<sup>1</sup> Lecture Table Experiment.

cury remains unchanged in the limb *ac* but it sinks gradually in the limb *de*. As soon as the level of the mercury has almost sunk to the tap *e*, close the tap *e*. The eudiometer is then ready for use.

(2) *To prepare the oxygen*.—Heat some powdered potassium chlorate in a small glass retort and after allowing the oxygen evolved at first to escape, collect the rest in a small gas-holder over water ready for use. The "oxygen" supplied for use with magic lanterns is not pure enough.

(3) *To introduce the hydrogen*.—The hydrogen must be prepared in the usual way from zinc and sulphuric acid, taking great care to avoid any admixture of air. For this purpose Kipp's apparatus (p. 333) is particularly suitable. After all air has been expelled from the apparatus and the leading tube, connect this tube with the tap *a* of the eudiometer, and at once open that tap. Hydrogen passes in and the mercury sinks down *ac* and rises up *ed*. As soon as the level of the mercury is nearly the same in both arms, close the tap, and at once disconnect the supply of hydrogen. Make the level of the mercury the same in both limbs, either by adding mercury to the limb *de*, or by running some out of the tap *e*, according as it is necessary. Read the volume occupied by the hydrogen. In an experiment it was found to be 54 c.c.

(4) *To introduce the oxygen*.—Attach a rubber tube to the gas-holder containing the oxygen. Run out a little oxygen, so as to expel all air from the tube, and then connect with *a*. Open *a*, and introduce about 10 c.c. of oxygen, and then close *a*. The level of the mercury must again be made the same in both limbs. Read the volume of gas enclosed. In the same experiment the total volume was found to be 66 c.c. Then, 12 c.c. oxygen had been added.

(5) *To fire the mixed gases*.—Effect this by connecting the wires at *b* with a Ruhmkorff coil and a bichromate cell.<sup>1</sup> The mercury at once rises part way up *ac*. Pour some mercury down *de* so as to make its level the same again in both limbs. Read the volume occupied by the gas. In the experiment quoted it was 29.5 c.c. Then, the contraction was  $66 - 29.5$  or 36.5 c.c. Why did a contraction take place? Because the

<sup>1</sup> See Appendix, p. 333.



steam formed by the union of the oxygen with some of the hydrogen was condensed, and occupied a negligible volume.

(6) *To show that the gas left is hydrogen.*—Pour some mercury into *d*. Open the tap *a*, and apply a flame to the issuing stream of gas. It burns with a pale flame. Then it is hydrogen, and there was more than sufficient hydrogen to unite with all the oxygen.

(7) *Conclusion.* Hence, in the experiment above referred to, the volume of hydrogen combined in the water formed was

$$36.5 - 12 \text{ or } 24.5 \text{ c.c.}$$

Now this was almost exactly twice the volume of the oxygen with which the hydrogen combined.

Therefore, this experiment leads to the conclusion that *two volumes of hydrogen and one volume of oxygen, are combined in forming water.*<sup>1</sup>

In 1808 Gay Lussac carried the investigation a stage further, and measured the volume of steam formed on the combination of hydrogen and oxygen.

EXPT.<sup>2</sup> 2. **To find the Volume of steam formed when Two volumes Hydrogen combine with One Volume Oxygen.**—Since it is desired to measure the volume of steam produced on exploding a mixture of two volumes hydrogen and one volume oxygen, it is essential to prevent the condensation of the steam. This is secured by surrounding the limb *a* of a eudiometer tube containing the mixed gases with a wider tube *b*, through which the hot vapour of boiling amyl alcohol can be passed. The alcohol, which boils at 132° C., is

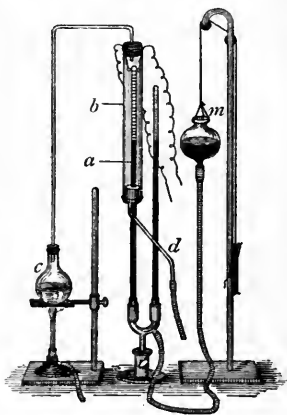


FIG. 132.

<sup>1</sup> Morley has recently (1895) determined the ratio of the volume of hydrogen and oxygen to be 2.00269 : 1.

<sup>2</sup> Lecture Table Experiment.

heated in the flask *c*. The vapour after passing through the jacket is led away by the tube *d*.

The mixture of two volumes hydrogen and one volume oxygen are introduced into the eudiometer tube *a*. When the temperature of the jacket is constant at  $132^{\circ}$ , bring the mercury to the same height in both limbs. Read off exactly the volume of the mixed gases. Next reduce the pressure on the gases by lowering the reservoir of mercury *m*. Now fire the mixture by an electric spark. Raise the reservoir again until the level of the mercury is the same in the two tubes. Read the volume. It is almost exactly two-thirds of that of the mixture of hydrogen and oxygen.

Hence, *when 2 volumes hydrogen and 1 volume oxygen combine, 2 volumes of steam or gaseous water are formed*

## THE COMPOSITION OF WATER BY WEIGHT.

(1) ANALYSIS.—Lavoisier was the first to carry out an analysis of water, by passing a known amount of water in the form of steam over red-hot iron and determining the amount of hydrogen set free (*see* page 223). Accurate results, however, could not be obtained by such a method.

(2) SYNTHESIS.—Water is produced when hydrogen is burnt or exploded with air. Then, if the water formed were weighed, and also *either* the oxygen *or* hydrogen which combined together, we should be able to arrive at its composition by weight. But it would be difficult to weigh either the oxygen or the hydrogen accurately.

Can water be synthesised in any other way? Is it possible that hydrogen can obtain the necessary oxygen from any other source than the air or oxygen gas itself? Can hydrogen obtain the oxygen from any of the oxides we know? To answer this we must appeal to experiment.

EXPT. 3. To find whether Water is formed by the action of Hydrogen on Red-hot Oxide of Copper.—Place some *dry* copper oxide in a hard glass tube *ab*, about 12 inches in length (Fig. 106). Place this over a Ramsay burner. Pass hydrogen through the tube, drying the gas first by bubbling it through strong sulphuric acid in a flask *c*. After sweeping out the air with the hydrogen, light the burner. Notice that a dew

is very soon deposited on the cooler portion of the tube near *b*, and that the surface of the oxide has become red in colour. Then the copper has in part lost its oxygen and the dew must be water formed by this oxygen uniting with some of the hydrogen gas.

We shall now be able to find the composition by weight of water, if we can weigh the whole of the water formed in the above experiment, and also the loss of oxygen from the copper oxide.

**EXPT. 4. To find the Composition of Water by Weight.**—Take a U-tube *ab*, of the form shown in Fig. 133,

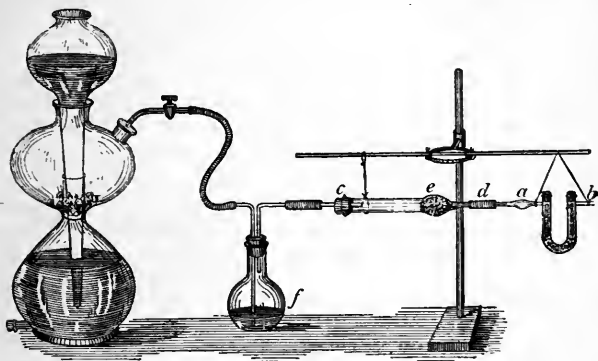


FIG. 133.

with two side tubes, on one of which a bulb *a* is blown. Nearly fill it with small lumps of dry calcium chloride. Close the two mouths with tightly fitting rubber stoppers, or better still, seal them in the blowpipe flame. Put a cap of rubber tubing closed by a small glass rod over the end of each side tube.

Dry some copper oxide (granular or in wire) by heating it strongly in a porcelain basin. Obtain a tube *cd* of hard glass on which a bulb *e* has been blown, and drawn out at one end.<sup>1</sup> Plug one end of the bulb loosely with a little dry asbestos. Fill the bulb with some of the copper oxide while still warm. Introduce another plug of asbestos. Close each end of the tube

<sup>1</sup> These are sold as reduction tubes.

with stoppers or caps. Take a hydrogen-Kipp apparatus. Connect it with a flask *f* containing some pure strong sulphuric acid, so that the hydrogen can be bubbled through and dried by the acid.

Weigh the tube *cd* (unstoppered) accurately. Connect it by a cork and tubing with the flask, supporting it by a clamp. Cause a slow current of hydrogen to pass through; it should be possible to count the bubbles as the gas passes through the acid in *f*. Weigh the U-tube (unstoppered) accurately. Fix it to the end *d* of the tube by a small piece of rubber tubing; the two glass tubes should be brought close together. Now heat the bulb *e* with a Bunsen burner, at first gently and then more strongly. The water which is formed is carried forward and is for the most part condensed in the little bulb *a* on the U-tube. If any dew is found condensed in the hard glass tube between *e* and *d*, the tube must be gently warmed. When 1 to 2 grams water appear to have collected in the bulb *a*, gradually cease heating *e*. Allow the bulb to cool in the current of hydrogen during half an hour. Detach the Kipp and flask. Suck air through the tube *cd* and the U-tube for a few seconds. Why is this necessary? Weigh the U-tube, and also the tube *cd* accurately. Here is the result of an experiment:—

Gain in weight of U-tube, *i.e.* weight of water formed = 1.191 grm.

Loss in weight of copper oxide, *i.e.* weight of oxygen lost = 1.067 grm.

Difference, or weight of hydrogen in water = .124 grm.

$$\begin{aligned}\text{Hence, } \frac{\text{weight of oxygen}}{\text{weight of hydrogen}} &= \frac{1.067}{.124} \\ &= \frac{8.6}{1}\end{aligned}$$

The experiment should be repeated several times. Do you obtain a similar result? The above result agrees fairly well with that which is established by more accurate experiments, *viz.* that *water consists of 8 parts by weight of oxygen for every 1 part by weight of hydrogen.*

**Reduction.**—Heated copper oxide, over which hydrogen is passed, loses oxygen and metallic copper results. The copper oxide is said to be *reduced*. *Reduction*, then, is the reverse of oxidation, and consists in the removal of oxygen from a substance

which contains it. Note that at the same time that the copper oxide is reduced the hydrogen passing over it is oxidised.

In a former experiment, see p. 222, steam was passed over red-hot iron, with the result that hydrogen was liberated and oxide of iron formed. In that case the vapour of water was reduced, while the metal was oxidised.

In both of these cases the two processes of oxidation and reduction took place simultaneously.

### PROBLEMS

1. Find whether (*a*) water is produced, (*b*) the oxide is reduced to metal by the action of hydrogen upon heated oxide of iron, red-lead, litharge, oxide of tin, manganese dioxide, oxide of magnesium, quick-lime, zinc oxide, bismuth oxide, cobalt oxide, antimony oxide.

2. Find the effect of passing a stream of coal gas over hot copper oxide, red-lead, and oxide of iron.

## CHAPTER XXXII

### ACIDS AND ALKALIES : THEIR COMBINING PROPORTIONS

#### FIXED OR VARIABLE PROPORTIONS?

We found a short time ago that a salt was produced when an acid was added to an alkali until the solution became neutral. Is there possibly any definite relation between the weights of acid and alkali which are concerned in producing any given salt? In other words, does a definite quantity of an acid combine with a fixed quantity of an alkali, or can it combine with any quantity that is mixed with it? This is a most important question and we must endeavour to work out the answer to it.

The best known of all "salts" is common salt, so we cannot do better than try to find whether it is formed from fixed or variable proportions of hydrochloric acid and of caustic soda. We can test this, in more than one way, by the three following experiments.

**EXPT. 1. To find how many c.c. of a given Hydrochloric Acid Solution are required to neutralise 1 c.c. of a given Caustic Soda Solution.**—The apparatus required will be a burette and stand, a 10 c.c. pipette, and 3 small flasks. Before beginning the experiment the student should read again the precautions to be taken in working with burettes and pipettes (*see* pp. 30, 31).

Exactly 10 c.c. of caustic soda solution is to be placed in each of the flasks. Wash the flasks out with water. Rinse the pipette with some of the soda. [Would it be a mistake to leave

it dry or wet with water?] Run 10 c.c. of caustic soda into each flask. To each flask add a small quantity of litmus, enough to colour them blue. Will the quantity of litmus used affect the result? If you don't know, perhaps it would be better to add equal quantities of litmus.

Rinse the burette first with water and then with some of the hydrochloric acid. [Why?] Fill the burette with acid. Run out a little through the tap to sweep away the air bubble which sometimes lurks there. Take the reading of the burette. Run acid, about 1 c.c. at a time, into one of the flasks, shaking the flask after each addition. When the litmus begins to redden run more slowly, say 1 c.c. at a time. Stop, if you can, when the litmus is exactly neutral. Probably you will overshoot the mark, and leave it red with excess of acid. Read the burette. Suppose the readings have been 3.5 c.c. and 15.7 c.c., with the difference 12.2 c.c. Now, this answer is wrong; it is too great. But if you have been adding acid 1 c.c. at a time you know that the answer is between 12.1 and 12.2 c.c. Repeat the experiment with flask no. 2. You know that more than 12 c.c. are wanted. To save time begin by running in 12 c.c. boldly, watching the burette scale. Shake the flask. The liquid is still blue. Now proceed very carefully, adding only *single drops*. Probably you will again go too far, and leave the litmus red instead of neutral. Repeat this experiment again and again—it is excellent practice—until you have 3 flasks all exactly neutral. Take readings to .01 c.c.

Your results will then stand—

1st trial ; from	3.5	to	15.7	=	[12.2]	red	W(rong).
2nd	15.7	to	27.8	=	[12.1]	red	W.
3rd	27.8	to	39.85	=	12.05	neutral	R(ight).
4th	1.23	to	13.26	=	12.03	neutral	R.
5th	13.26	to	25.34	=	[12.08]	red	W.
6th	25.34	to	37.38	=	12.04	neutral	R.

You find, then, that a definite volume of caustic soda is neutralised in each case by almost exactly the same volume of hydrochloric acid:

Can you average these results?

$$\begin{array}{r} 6)72.50 \\ 12.08 \end{array}$$

You obtain 12.08 :

But you have included the wrong results in your average. Those in brackets should be rejected, and only those marked R reckoned. Hence, averaging,

					$3\overline{)36\cdot12}$				
	10 c.c. of caustic soda required				12'04 c.c. hydrochloric acid.				
Then,									
	1 c.c.	,	,	,	1'204 c.c.	,	,	,	

### PROBLEM

Repeat the experiment placing the soda solution in the burette, and using the acid in the pipette. Do you get the same answer? If not, why not?

Wash the burette very carefully (as your custom should be) after this experiment for fear the soda should fix the tap tight.

**EXPT. 2.<sup>1</sup> To find the Weight of Common Salt formed on neutralising 20 c.c. of Caustic Soda with Hydrochloric Acid.**—Weigh 3 light basins. Measure into each 20 c.c. of caustic soda. Measure from a burette the amount of hydrochloric acid necessary to neutralise the caustic soda in one basin. Measure *twice* this volume of acid into the second basin, and three times the volume into the third basin. Evaporate the contents of each basin to complete dryness. Find the increase in weight of each basin.

Do you not find that the same amount of salt is produced in each case? Then the extra amount of acid added to the second and third basins, beyond that necessary to render the solutions neutral, did not combine with the caustic soda, but was evaporated.

**EXPT. 3.<sup>1</sup> To find whether the Weight of Salt formed and the Volume of Acid used in neutralising Caustic Soda is proportional to the Quantity of Soda taken.**—Suppose we measure out volumes of caustic soda solution in the ratio 1 : 2 : 3, and neutralise each with hydrochloric acid. Then, if the alkali and acid act, as they appear to do, in fixed proportions, the volumes of acid added should also be in the

<sup>1</sup> Different members of a class should be set to carry out the several experiments at the same time. The results should then be all discussed together.



ratio 1 : 2 : 3. Moreover, if we determine the weights of salt formed, these should be in the same ratio.

Proceed thus :—Weigh three light basins. Measure into them 10, 20, and 30 c.c. respectively of dilute caustic soda solution. Add to each a few drops of litmus solution. Measure with a burette the volume of dilute hydrochloric acid which must be run into each so as to neutralise the solutions. Evaporate the solutions to dryness. Find the increase in weight of the basins. State your results in the way that the results of some actual experiments are recorded here :—

Vol. of Caustic Soda in c.c.	Vol. of Hydrochloric Acid in c.c.		Weight of Salt in grams.	
10	12·8	$\div 1 = 12\cdot8$	1·65	$\div 1 = 1\cdot65$
20	25·4	$\div 2 = 12\cdot7$	3·20	$\div 2 = 1\cdot60$
30	38·5	$\div 3 = 12\cdot8$	4·81	$\div 3 = 1\cdot60$

The results of experiments 1, 2, and 3 make it certain that common salt is formed from *definite* and *fixed* quantities of hydrochloric acid and caustic soda, and not from haphazard and variable amounts. But you also must try whether *your own* experiments lead to the same conclusion.<sup>1</sup>

You must notice that we have not proved that salt is the only product, or that the weight of salt produced in any case is equal to the weight of caustic soda and the weight of the hydrochloric acid which act on one another. To assure ourselves of this we should need to weigh the amounts of soda and acid which take part in the action. This is not readily done, so we must be careful not to assume that which we have not proved.

**Definition of equivalent to.**—*The weight of an acid and the weight of an alkali which exactly neutralise each other are said to be equivalent to each other.*

<sup>1</sup> Remember that since the relative strengths of the caustic soda and hydrochloric acid solutions you employ are probably not the same as those used in the experiments quoted you will consequently not obtain the same numbers.

## EQUIVALENT QUANTITIES OF ACIDS AND ALKALIES.

We have just established the fact that caustic soda combines with hydrochloric acid in a fixed proportion. Now caustic potash resembles caustic soda very closely both in appearance and in properties. Supposing then we dissolve the same weights of caustic soda and of caustic potash in equal volumes of water, we may naturally ask whether the same volumes of the two alkaline solutions will neutralise equal weights of hydrochloric acid, or of any other acid? In other words, are the same weights of caustic soda and caustic potash equivalent to equal weights of acids?

This is a problem which is well worth solving. It was investigated and solved in Berlin towards the end of last century, by a German chemist named *Richter*, and we cannot do better than endeavour to work it out for ourselves.

Now in order to study this problem we must work with solutions of acids and alkalies of known strength. These then must be prepared in the first instance.

### EXPT. 4. To prepare Standard Solutions of Acids and Alkalies.

(1) *Caustic Soda*.<sup>1</sup>—Counterpoise a dry beaker. Place 30 grams on the right-hand pan. *Quickly* add sufficient solid caustic soda to the beaker to just restore equilibrium. Pour water into the beaker and stir with a glass rod until the soda has dissolved. Carefully pour the solution down the glass rod into a litre flask. Rinse out the beaker with a little water, and pour into the flask. Rinse three times more. Fill up the flask to the mark with water.

Then,

$$1 \text{ c.c.} = \cdot 05 \text{ gram caustic soda.}$$

Transfer the solution to a Winchester pint bottle. Write this value on a gummed label, and fix it on the bottle.

(2) *Caustic Potash*.<sup>1</sup>—Prepare in the same way a solution of solid caustic potash (in sticks), such that

$$1 \text{ c.c.} = \cdot 05 \text{ gram caustic potash.}$$

<sup>1</sup> Solid caustic soda and caustic potash are rarely quite pure. Hence, since the amount of impurity present is not always the same, the numbers obtained will probably not be the same as those quoted in this chapter.



## PROBLEMS

1. Find the weight of caustic potash neutralised by 100 grams of sulphuric acid.
2. Find the weight of (a) caustic soda, (b) caustic potash neutralised by 100 grams of hydrochloric acid.
3. Find the weight of (a) caustic soda, (b) caustic potash neutralised by 100 grams of oxalic acid.

Compare your results with the following :—

TABLE I

	Weights of Alkalies equivalent to 100 grams of acid.	
	Caustic Soda.	Caustic Potash.
Sulphuric acid ... ..	84·25	128·45
Oxalic acid ... ..	90·65	138·55
Hydrochloric acid ...	119·80	182·20

**Discussion of Results.**—Your experiments will probably lead to numbers which are nearly the same as the above. If so we are led to the following conclusions :—

(1) The weight of caustic soda which neutralises 100 grams of any particular acid is not equal to the weight of caustic potash which neutralises 100 grams of the same acid.

(2) The weights of caustic soda which severally neutralise 100 grams of *different* acids are not identical. A similar statement is true of caustic potash.

(3) The weights of caustic soda and of caustic potash which are equivalent to a given weight of an acid seem in each case to be in about the same proportion to one another.

Let us test this third conclusion more closely, and calculate in each case the *ratio* between the equivalent amounts of the two alkalies :—

$$\frac{84\cdot25}{128\cdot45} = \cdot656.$$

$$\frac{90\cdot65}{138\cdot55} = \cdot654.$$

$$\frac{119\cdot80}{182\cdot20} = \cdot657.$$

Then, the  $\frac{\text{soda}}{\text{potash}}$  ratio is the same whether we use sulphuric, hydrochloric or oxalic acid as the means of comparison. This surely is a very remarkable fact.

(4) In 1802 a German chemist, *G. E. Fischer*, pointed out that the results of such experiments as we have been making could be so calculated that all the quantities should be equivalent to one another. For instance, the first line in Table I. shows that 100 grams of sulphuric acid are equivalent to

$$\begin{array}{l} 84.25 \text{ grm. caustic soda, and to} \\ 128.45 \text{ grm. caustic potash} \quad \dots \quad \dots \quad \dots \quad \text{(i).} \end{array}$$

Now, if we multiply each quantity stated in the second line of Table I. by  $\frac{84.25}{90.65}$ , we shall make the quantity of the soda the same as that in the first line of the Table. The second line will then become—

$$100 \times \frac{84.25}{90.65} \text{ grms. of oxalic acid are equivalent to}$$

$$90.65 \times \frac{84.25}{90.65} \text{ grm. soda,}$$

$$\text{and to } 138.55 \times \frac{84.25}{90.65} \text{ grm. potash,}$$

or, 93 grms. of oxalic acid are equivalent to

$$\begin{array}{l} 84.25 \text{ grm. soda, and to} \\ 128.9 \text{ grm. potash} \quad \dots \quad \dots \quad \dots \quad \dots \quad \dots \quad \text{(ii).} \end{array}$$

Similarly, if we multiply each quantity stated in the last line of Table I. by  $\frac{84.25}{119.80}$ , we shall make the quantity of the soda also the same as that in the first line of the Table. The last line will then become—

$$100 \times \frac{84.25}{119.80} \text{ grm. hydrochloric acid are equivalent to}$$

$$119.80 \times \frac{84.25}{119.80} \text{ grm. soda,}$$

$$\text{and to } 182.20 \times \frac{84.25}{119.80} \text{ grm. potash,}$$

or, 70.33 grm. hydrochloric acid are equivalent to

$$\begin{array}{l} 84.25 \text{ grm. soda, and to} \\ 128.1 \text{ grm. potash} \quad \dots \quad \dots \quad \dots \quad \dots \quad \dots \quad \text{(iii).} \end{array}$$

Tabulating the statements (i), (ii), and (iii), we can form the following Table of equivalent quantities, expressed in grams.

TABLE II

ACIDS.	ALKALIES.	
	Caustic Soda.	Caustic Potash.
Sulphuric acid ... 100	84.25	128.5
Oxalic acid ... 93	84.25	128.9
Hydrochloric acid 70.3	84.25	128.1
		mean = 128.5

The weights of acids given in the first column are equivalent to the same weight of caustic soda. They are, therefore, equivalent to one another, for *quantities that are equivalent to the same quantity are equivalent to one another.*

But the numbers in the third column show that the weights of acids given in the first column are also equivalent to almost exactly the same weights of caustic potash. The difference between them, which is only in the decimal place, is probably simply due to small errors in the experiments.

Hence, Table II. can be further simplified to Table III., which expresses weights in grams of the acids and alkalies we have studied, such that any one is equivalent to any other, whether acid or alkali.

TABLE III

ACIDS.	ALKALIES.
Sulphuric acid 100	Caustic soda 84.25
Oxalic acid 93	Caustic potash 128.5
Hydrochloric acid 70.3	

We seem, through our study of the proportions in which acids and alkalies neutralise one another, to be approaching some

great principle, which underlies all the facts we have met with. What can there be in the nature of acids and alkalies which leads to the definiteness of combination and of neutralisation which we have found?

### PROBLEMS AND EXERCISES

#### Neutralisation, &c.

1. Given two solutions of sulphuric acid A and B (of which A contains 49 grams per litre), and a solution of alkali, find how many grams of acid per litre are contained in B.

2. C contains 49 grams of sulphuric acid; and D contains 50 grams per litre. How many c.c. of D must be taken to make after dilution a litre of the same strength as C? Dilute accordingly. Then, compare by experiment the strength of C with that of diluted D. Express your answer as a *ratio*:—Grams per litre in diluted D  $\div$  grams per litre in C.

3. X contains 50 grams, and Y contains 46 grams of caustic soda per litre. How many c.c. of water must be added to one litre of X to dilute it to the same strength as Y? Dilute accordingly, and compare by experiment the strengths of diluted X with that of Y.

4. Given a solution of sulphuric acid and some solid caustic soda. Make a strong solution of the soda, and find out in what proportion to add water so as to dilute it to a strength *equivalent* to that of the sulphuric. Add *nearly as much* water as you expect necessary. Find by experiment and by calculation how much more is required. Dilute accordingly.

5. Dry some anhydrous potassium carbonate by heating it strongly in a basin. Weigh out 1 gram. Dissolve it in hot water. Add litmus. Run in standard sulphuric acid. Boil to get rid of carbonic acid. Find what weight of acid is required to neutralise the solution.

6. Repeat No. 5, using hydrochloric acid in place of sulphuric acid.

7. Repeat No. 5, using dry sodium carbonate in place of potassium carbonate.

8. Find how many c.c. of the solution P are required to neutralise 10 c.c. of the solution Q (*a*) using P, (*b*) using Q in the burette.

9. Find how many c.c. of the acid solution R are equivalent to 10 c.c. of the acid solution S.

10. Find how many c.c. of the acid solution K are equivalent to 10 c.c. of the alkaline solution L, (*a*) using methyl-orange, (*b*) using phenolphthalein as indicators.

## CHAPTER XXXIII

### EQUIVALENT WEIGHTS OF THE METALS

**THE REPLACEMENT OF ONE METAL BY ANOTHER METAL.**—It was known long ago that when a piece of iron was placed in a solution of copper sulphate a red deposit of copper was quickly formed upon the surface of the iron, and it was generally believed that the iron was converted into copper.

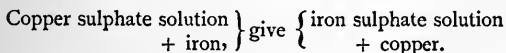
**EXPT. I. To find what occurs when Iron is placed in a Solution of Copper Sulphate.**—Place about 50 c.c. of a rather dilute solution of copper sulphate in a beaker. Warm and add some pieces of clean bright iron (turnings, wire, or nails). Notice the red deposit which is instantly formed on the iron. The blue colour becomes fainter. Go on adding iron until at last the liquid ceases to stain the iron. The blue colour finally disappears entirely. Filter the nearly colourless liquid from the red deposit and the iron into a porcelain basin.

The red deposit is evidently copper. Place a small portion in a test-tube, and test it with nitric acid. Yes, it is copper. Is anything left dissolved in the water? Test a solution of copper sulphate with litmus paper. It is slightly acid. Test also the solution obtained above. It also is slightly acid. Does it now only contain free sulphuric acid, since the copper of the copper sulphate has been precipitated? To decide this boil down the colourless solution. A red-brown precipitate is gradually formed. Is this copper? Test a portion with nitric acid. It is not copper. Is it then a substance containing iron, perhaps iron



rust? How can we test this? Suppose we try the effect of heating it in hydrogen. What will the effect be if it is iron rust? Proceed by drying some of the red precipitate, and heating it in a current of hydrogen. Do you not find it to behave just like the rust or oxide of iron (*see* p. 286, Expt. 3)?

So *iron* has passed into solution, and *replaced the copper*. The action therefore is :—



### PROBLEMS

(1) Find whether zinc, magnesium, tin, and lead can precipitate copper from a solution of copper sulphate.

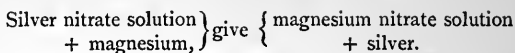
(2) Find whether copper can precipitate iron from a solution of iron sulphate.

It seems improbable that copper is the only metal which can be displaced from its solution in this way. Suppose we try the effect of metals upon a solution of silver nitrate.

**EXPT. 2. To find what occurs when Magnesium is placed in a Solution of Silver Nitrate.**—Dissolve 2 or 3 crystals of silver nitrate in about 20 c.c. of water in a test-tube. Add a small piece of clean bright magnesium ribbon. A gray deposit begins to form on the magnesium. Shake. The powdery deposit falls to the bottom. The magnesium seems to become thinner, and finally disappears. Add another small piece. Proceed in this way *until no further deposit is seen to form*.

What is the deposit? What became of the first piece of magnesium? Pour the liquid and the deposit on to a filter, collecting the filtered liquid in a basin. Evaporate the liquid to dryness. A white residue with a bitter taste is obtained. It is not silver nitrate but magnesium nitrate. Meantime wash the deposit on the filter by pouring hot water on several times. Dry it carefully over a flame (*see* p. 123). It becomes lighter in colour. Rub it with a bright knife-blade. It acquires a silvery lustre. So the gray deposit is evidently silver.

The action therefore is :—



### PROBLEMS

1. Find the effect of iron, zinc, and copper upon a solution of silver nitrate.
2. Find whether a silver coin will precipitate solutions of the sulphates of iron, zinc, and copper.
3. Find the effect of iron, magnesium, and copper upon a solution of lead nitrate.
4. Try the effect of zinc upon a solution of lead acetate by hanging a piece of granulated zinc in a beaker containing a saturated solution ; put it aside for a day.
5. Try to draw up a list of the metals experimented with in such an order that each one will precipitate those below it in the list.

We can no longer doubt that one metal can frequently precipitate another metal from its solution. The precipitating metal at the same time is always dissolved to a greater or less extent, so that it is a case of one metal replacing another. We have been discovering for ourselves what *Bergman*, a famous Swedish chemist, first discovered more than 100 years ago. Lavoisier, also, studied this problem in 1782.<sup>1</sup>

Is there any connection between the weight of metal precipitated and the weight of the metal which passes into solution? For instance, will a definite amount of iron replace an unlimited quantity of copper? It does not seem likely. Careful measurements, however, must be made to settle the question.

**EXPT. 3. To find the Weight of Iron which will replace 100 Grams of Copper.**—Measure about 100 c.c. of copper sulphate solution into a beaker. Heat it nearly to boiling. Clean some iron pianoforte wire with emery paper. Cut it into short pieces. Weigh out carefully between .3 and 1 gram of the wire. Place it in the hot copper solution. Stir continually with a glass rod, until the iron wire is entirely dissolved.

Place a filter paper on one pan of a balance. Counterpoise it exactly with another paper on the other pan. Fold the two filters,

<sup>1</sup> Lavoisier's *Works*, vol. ii. 528-544.

place one inside the other and then place them in a funnel. Carefully pour the contents of the beaker while still warm on to the filter, collecting the filtrate in another beaker. Wash *all* the precipitated copper on to the filter. Wash it several times by filling the paper with warm water. It must now be dried. Wash the paper in the same way twice with methylated spirit. Take the filters out of the funnel. Open them carefully. Place them on three or four other filter papers. Place the pile on wire gauze six inches above a flame, as shown in Fig. 70, until quite dry. Allow to cool.

Then remove the top paper (bearing the copper), and place it on the balance pan. Place the second paper on the other pan. Find the increase in weight.

Thus,

·28 gram of copper was replaced by ·25 gram of iron.

Then,

100 grm. copper would be replaced by  $\cdot 25 \times \frac{100}{\cdot 28}$  or 89·2 grm. iron.

Now it is most important to find whether other experiments, carefully carried out, *always* lead to the same result. Repeat the experiment with varying amounts of iron. Calculate in each case the weight of iron which would, according to the experiment, replace 100 grams of copper.

Here are some results expressed in grams :—

Weight of Copper replaced.	Weight of Iron taken.	Calculated Weight of Iron replacing 100 grams of Copper.
·28	·25	89·2
·325	·288	88·6
·39	·344	88·2
·392	·352	89·8
·42	·374	89·0
		mean 89·0

Hence the actual weight of copper replaced is increased at the same time as the weight of iron used to precipitate it. But in each case if calculated per 100 grams of copper, 100 grams

of copper are replaced by very nearly the same weight of iron (about 89 grams). It seems as though the slight differences are due to small errors in making the experiments. Probably, therefore, if we could in each case make the experiments without any error, always weighing quite correctly, never losing any copper, washing thoroughly and drying completely, then we should always find that 100 grams copper are replaced by exactly the same amount of iron. Probably the number 89.0, which is the mean result of the experiments, expresses fairly accurately the true weight of copper replaced by 100 grams of iron. But you must find whether *your own* experiments also lead to this result.

**DEFINITION.**—*The weight of one metal, and the weight of another metal which can replace it are said to be equivalent to one another.*

Will the same weights of other metals, for example 89 grams of zinc and magnesium, also replace or be equivalent to 100 grams of copper? This must be answered by experiment.

#### PROBLEMS

Find what weight of (1) zinc, (2) magnesium will replace 100 grams of copper.

(Take .2 to .5 gram of zinc foil,<sup>1</sup> and .1 to .4 gram of magnesium. The metal must be rubbed bright with emery paper. The zinc must be placed in the *cold* copper solution. It will probably be necessary to warm the solution slightly in the case of magnesium in order to dissolve the whole of it.)

**Results.**—The results of a number of experiments gave the following mean numbers :—

100 grams of copper were replaced by 38 grams of magnesium.				
“	“	“	89	“ iron.
“	“	“	103	“ zinc.

It certainly appears that *definite* and *fixed* weights of metals can replace, or are equivalent to, a fixed weight of another metal. These weights, moreover, appear to differ between themselves.

<sup>1</sup> Pure distilled zinc foil can be obtained at 6s. a lb.

If we multiply each of these numbers by  $\frac{100}{38}$  we obtain the following :—

*Weights of the Metals equivalent to 100 Grams of Magnesium.*

100	grams of magnesium.
234	,, iron.
263	,, copper
271	,, zinc.

In order to make more certain of this definiteness of replacement it will be well to examine the replacement of some other metal. Suppose we study the replacement of silver by other metals from a solution of silver nitrate.

**EXPT. 4. To find the Weight of Magnesium which will replace 100 Grams of Silver.**—Weigh out roughly upon a piece of paper two to three grams of crystals of silver nitrate. Transfer the crystals to a beaker. Add about 100 c.c. of water, and stir with a glass rod until the crystals are dissolved. Weigh out accurately from 0.1 to 0.15 gram of clean magnesium ribbon. Place the magnesium in the solution. Stir so as to shake off the gray deposit of silver which forms upon the ribbon, until the whole of the magnesium disappears. Pour the whole of the contents of the beaker upon a filter paper, and proceed exactly as described in the last experiment. (Keep the filtrate in a bottle labelled "Silver residues.")

In an actual experiment

1.47 grams of silver were replaced by 0.161 gram of magnesium.

Then,

100 grams of silver would be replaced by 10.9 grams of magnesium.<sup>1</sup>

### PROBLEMS

Find what weight of (1) zinc, (2) copper, will replace 100 grams of silver. Take about 0.2 gram of metal (pure foil) in each case. (The zinc should be placed in a *cold* solution of silver nitrate; warm towards the end if the zinc has not passed entirely into solution. The copper must be placed in a *hot* solution of silver nitrate.)

<sup>1</sup> Lower values than this will be obtained sometimes.

**Results.**—The results of a number of experiments gave the following mean numbers :—

100	grams	silver	were	replaced	by	11	grams	of	magnesium.
29.5	grams	copper	replaced	100	grams	silver			
30.5	grams	zinc	replaced	100	grams	silver			

If we multiply each of these numbers by  $\frac{100}{11}$  we obtain the following :—

*Weights of the Metals equivalent to 100 grams of Magnesium.*

100	grams	magnesium.
268	grams	copper.
277	grams	zinc.
909	grams	silver.

Compare these numbers with those obtained from the experiments with the copper sulphate solution. Nearly the same numbers are obtained for the weights of copper and zinc equivalent to 100 grams of magnesium.

### SUMMARY

Let us state clearly the several important conclusions we have arrived at :—

(1) One metal can frequently be precipitated by another metal from a solution of the first metal.

(2) The precipitating metal is at the same time dissolved.

(3) If a metal A can precipitate a metal B from its solution, then B will not precipitate A from its solution.

(4) A fixed weight of one metal is replaced by a fixed weight of another metal.

(5) 100 grams copper are replaced by different weights of magnesium, iron, and zinc.

(6) 100 grams of silver are replaced by different weights of magnesium, zinc, and copper.

(7) The weights of magnesium and zinc which respectively replace 100 grams of copper are different from the weights of these two metals which replace 100 grams of silver.

(8) The experiments with the copper sulphate solution and those with the silver nitrate solution lead respectively to almost the same numbers for the weights of zinc and copper which are equivalent to 100 grams of magnesium.

As we have already said Bergman in Upsala first studied the precipitation of one metal by another. But neither he, nor Lavoisier in Paris, who also a little later took up the problem, could satisfactorily explain why a definite fixed amount of one metal could replace another metal, and only a definite fixed amount of this second metal. Can you suggest any reason *why* one metal can replace another in this definite way?

## THE FORMATION OF METALLIC OXIDES.

(1) THE CALCINATION OF METALS.—When studying Fire and Air we found that when many metals were strongly heated in the air they gained in weight. We concluded that they combined with the Fire-Air or oxygen of the atmospheric air, whereby a calx or oxide of the metal was formed. But we did not consider either of the following questions :—

Will a given weight of a metal take up or combine with an unlimited amount of oxygen, or only with a definite fixed amount?

Will equal weights of different metals all combine to the same extent with oxygen, or will some combine with more than others?

Here are two problems. Experiment alone can answer them. Let us study first the calcination of magnesium, as we have found that its calx, *i.e.* an oxide of the metal, is readily formed.

**EXPT. 5. To find the Weight of Magnesium Oxide formed on Calcining 100 Grams of Magnesium.**—Weigh a half ounce porcelain crucible with its lid. Clean some fine magnesium ribbon with sand paper. Weigh out accurately from 0.2 to 0.3 gram of the magnesium on to the pan bearing the crucible. Roll the ribbon into small compass, and place it in the crucible. Support the crucible (with the lid) upon a pipe-clay triangle on a tripod stand (Fig. 134). Then apply a flame gradually until the bottom of the crucible is red hot. This should take five minutes. Continue heating strongly for another five minutes, but lift the lid *very slightly* every half minute. Every care must be taken to avoid the escape of fumes. Allow to cool. Weigh.

Thus,

·228 gram magnesium gave ·380 gram of magnesium oxide.

According to this,

1 gram magnesium would give 1.66 gram of magnesium oxide.

Now it is very important to find out whether magnesium on calcination *always* leads to the same result. Here are the results of a number of other experiments, but you also must make further experiments, and find out whether *your results* support these or not :—

Weight of Magnesium.	Weight of Oxide of Magnesium.	Calculated Weight of Oxide of Magnesium from 100 grams of Magnesium.
.12	.196	163
.158	.265	167
.211	.351	166
.27	.438	162
.277	.464	167
.341	.573	168
		mean 166

Hence,

(1) the actual weight of oxide formed is increased at the same time as the weight of magnesium calcined.

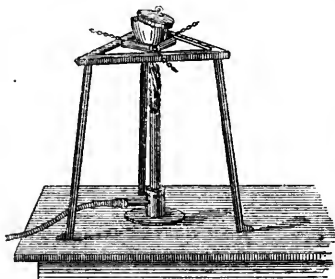


FIG. 134.

(2) In each case the weight of oxide formed is nearly proportional to the weight of magnesium calcined.

So far then as magnesium is concerned we can give an answer to the first question above, and say that : given weight of magnesium cannot combine with an unlimited weight of oxygen ; on the contrary, it combines with a *definite fixed* weight. For each experiment leads to the conclusion that 100 grams of magnesium can combine with no more than about 66 grams of oxygen.



Observe, however, that we have no right to answer the question in the same way for *other* metals.

Other metals gain in weight on calcination in the air, for example: zinc, tin, lead, bismuth, copper. Their *complete* oxidation by calcination in the air is however only slowly accomplished, and therefore this method of oxidation of the metals is not so suitable for our investigation.

### PROBLEMS

Study the increase in weight on calcination of (1) zinc, (2) tin, (3) lead, (4) copper, (5) iron, by aspirating (*a*) air, (*b*) oxygen over the metals placed in a porcelain boat within a piece of hard glass tubing heated by a Ramsay burner, or by a small furnace.

(2) THE OXIDATION OF METALS BY NITRIC ACID.—In a previous research (p. 211) we found that when a metal was dissolved in nitric acid, the gas given off was not hydrogen; for example, Nitrous Air, now called nitric oxide, was evolved when copper was treated with nitric acid. On evaporating the liquid, crystals of the nitrate of the metal were obtained. Suppose we try the effect of heat upon these crystals.

**EXPT. 6. To find the Effect of Heat upon Crystals of Copper Nitrate.**—Heat some crystals of copper nitrate in a test-tube. Steam is given off, and is condensed on the sides of the tube. Probably this is water of crystallisation. Then a red gas is evolved, and the residue blackens.

Take a rather larger sample, and heat it in a porcelain crucible, from which water can escape freely. Heat it, until there is no further change. A black substance remains. This must contain the copper of the copper nitrate. What substance containing copper is black? The oxide. Is this the oxide? Compare it with a sample of oxide. Find what happens on heating the black substance in a current of dry hydrogen. Do you not find that water is formed, and red copper left? Then, copper oxide is formed on heating copper nitrate.

### PROBLEMS

1. Find whether the oxides of lead, iron, and zinc can be prepared by dissolving the metals in nitric acid, and heating the products.

2. Find whether the product of heating tin with nitric acid is an oxide or a nitrate.
3. Find whether an oxide is left on heating silver nitrate.

It is clear that oxides of many metals can be obtained through the action of nitric acid upon them. Is nitric acid, then, a store-house of oxygen? Let us try the effect of heat on nitric acid.

**EXPT. 7. To show that Nitric Acid contains Oxygen.**

—Support a churchwarden tobacco pipe over a Ramsay burner, with the mouthpiece dipping into a pneumatic trough underneath a jar full of water (Fig. 135.)

When the portion of the stem in the burner is thoroughly heated, pour a few drops of strong nitric acid into the bowl.

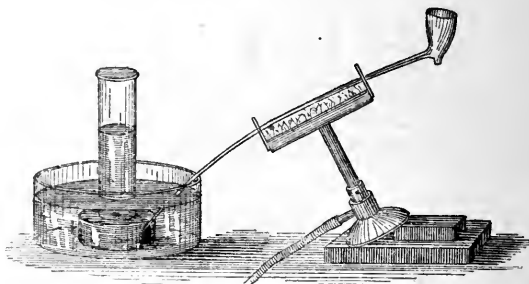


FIG. 135.

Reddish bubbles of gas pass up through the water in the jar. But the gas collected is colourless. Test the gas with a glowing splinter of wood. It is found to be oxygen.

Now if magnesium oxide is formed by dissolving magnesium in nitric acid and heating the product, may we expect to find the same weight of oxide formed by this method from 100 grams of the metal as by calcination in the air?

**EXPT. 8. To find the Weight of Magnesium Oxide formed on oxidising the Metal with Nitric Acid.**—Support a porcelain crucible and lid upon a pipe-clay triangle. Heat it, and then allow it to cool. Weigh it. Weigh into it accurately '3 to '6 gram of clean magnesium wire or fine ribbon. Place the crucible on a sand-bath in a draught-cupboard. Cover the magnesium with a small inverted funnel (Fig. 136).

Lift the funnel and add about 10 drops of concentrated nitric acid. Replace the funnel. Heat the sand-bath until a vigorous action has begun, and then remove the flame. When the action is over add five more drops of nitric acid. Lift the funnel and wash any liquid upon it into the crucible with a fine stream of water (as little as possible) from a wash-bottle. Put the funnel aside. Remove the dish from the sand-bath and place it on a pipe-clay triangle upon the ring of a retort-stand. Place a small flame underneath, and evaporate carefully until quite dry. (Every care must be taken to avoid loss by spurting.) What is the white residue? Heat again, more strongly. Red fumes are evolved. Finally heat with the full flame for five minutes. Allow to cool. A white residue remains. Weigh. Heat again and weigh, until the crucible is constant in weight.



FIG. 136.

Thus,

75 gram of magnesium gave 1.24 grams of magnesium oxide.

Then,

100 grams of magnesium would give 166 grams of magnesium oxide.

Thus through the action of nitric acid upon magnesium we have obtained exactly the *same weight* of a white residue as the weight of white oxide of magnesium formed by calcining 100 grams of magnesium in the air. Nor can any difference be detected between the two samples of oxide. They must, therefore, be identical.

Find whether *your* experiments confirm this conclusion.

We have now good ground for the belief that many metals can be oxidised by dissolving them in nitric acid and heating the products. This can be quickly accomplished, and is a more ready method of obtaining the oxides of the metals than that of calcining them in the air, which is often tediously slow.

## PROBLEM

Find what weight of oxide can be prepared from 100 grams of each of the following metals by the action of nitric acid : (1) zinc, (2) copper, (3) iron, (4) lead. Take .5 to 1 gram of each metal.

**Results.**—Compare your results with the following numbers, which are the mean results of a number of experiments :—

100 grams of magnesium gave 166 grams of oxide.

„	iron	„	143	„
„	copper	„	125	„
„	zinc	„	124	„
„	lead	„	107	„

Suppose we calculate from the above numbers the weight of each metal with which 100 grams of oxygen are combined. For example :—

100 grams of magnesium gave 166 grams of magnesium oxide.

Thus,

66 grams of oxygen combined with 100 grams of magnesium.

Then,

1 gram of oxygen combined with  $\frac{100}{66}$  or 1.51 grams of magnesium.

or,

100 grams of oxygen combined with 151 grams of magnesium.

In this way the following numbers were obtained :—

100 grams of oxygen combined with 151 grams of magnesium.

„	„	„	232	„	iron.
„	„	„	400	„	copper.
„	„	„	417	„	zinc.
„	„	„	1428	„	lead.

## SUMMARY.

Let us state clearly what we have established with regard to the oxidation of metals with nitric acid.

(1) 100 grams of each metal give a definite fixed weight of its oxide.

(2) 100 grams of one metal give a different weight of oxide from the weight of oxide obtained from another metal.

(3) 100 grams of oxygen combine with a definite fixed weight of a metal.

(4) 100 grams of oxygen combine with different weights of the several metals.

**Equivalent Weights of the Metals.**—The different weights of the several metals which combine with 100 grams of oxygen are for that reason said to be *equivalent* to one another, viz. :—

151	grams of magnesium.
232	„ iron.
400	„ copper.
417	„ zinc.
1428	„ lead.

If we multiply each of these numbers by  $\frac{100}{151}$  we obtain the following :—

*Weights of the Metals equivalent to 100 Grams of Magnesium.*

100	grams of magnesium.
150	„ iron.
265	„ copper.
279	„ zinc.
945	„ lead.

How do these numbers agree with those obtained from the study of the replacement of one metal by another? Tabulate the several results.

	Weights of the Metals equivalent to 100 grams of Magnesium.		
	Method 1. Replacement of Copper.	Method 2. Replacement of Silver.	Method 3. Oxidation of Metals with Nitric Acid.
Iron ...	234	—	153
Copper ...	263	268	265
Zinc ...	271	277	279
Silver ...	—	909	—
Lead ...	—	—	945

The metals iron, copper, and zinc are the only metals the

weights of which, equivalent to 100 grams of magnesium, have been determined by more than one method.

Observe that when these weights are placed in order of magnitude the *order of the metals is the same* by whichever method the weights have been obtained.

Notice that the equivalent weights of copper and of zinc are nearly the same as determined by all three methods. The greatest difference between the numbers for copper is 5, while the difference between those for magnesium and copper, two different metals, is more than 160.

The numbers for iron, on the contrary, do not agree well. Is there any simple relation between the numbers 234 and 153? Yes, for

$$234 \times \frac{2}{3} = 156, \text{ i.e. about } 153$$

or

$$153 \times \frac{3}{2} = 229.5, \text{ i.e. about } 234.$$

It appears then that in the case of iron there is more than one definite weight which is equivalent to 100 grams of magnesium. These two weights appear, moreover, to be related to one another in a simple way, viz. as 3 : 2.

We must, at any rate, keep before us the possibility of *some* metals having more than one equivalent value—a possibility recognised by *Richter*, a German chemist, who did much very valuable work on this subject towards the end of last century.

#### EXAMPLES XVIII

##### Equivalents of the Metals

1. If .04 gram of magnesium gave 40 c.c. of hydrogen on solution in hydrochloric acid, and .937 gram of zinc gave 353 c.c. of hydrogen, calculate how many grams of zinc will yield the same volume of hydrogen as 100 grams of magnesium.

2. If 1.95 grams of tin gave 417 c.c. of hydrogen, and .975 gram of zinc gave 365 c.c. on solution in hydrochloric acid, find how many grams of tin are equivalent to 1 gram of zinc.

3. If .474 gram of magnesium gave 465 c.c. of hydrogen, find the number of grams (a) iron, (b) zinc, equivalent to 100 grams of magnesium, when

(a) .996 gram of iron gave 432 c.c. of hydrogen.

(b) .926 gram of zinc gave 340 c.c. of hydrogen.

## CHAPTER XXXIV

### THE TRANSFORMATION OF MATTER.—THE LAW OF DEFINITE PROPORTIONS.

ONE great fact which has stood out in every inquiry and research upon which we have entered is this, that matter can undergo change or transformation. If we consider our work more closely we shall find that many of these very numerous instances of the change of matter can be classed under three heads :—

**Decomposition.**—Not infrequently the effect of heat has been to produce two substances from one. Thus :—

Chalk = quicklime + carbonic acid gas.

Red lead = litharge + oxygen.

Water = hydrogen + oxygen.

Such a change is a *decomposition* or an *analysis*.

**Combination.**—Numerous instances of combination were met with in our study of Fire and Air. Thus :—

Lead + oxygen = litharge, or else red lead.

Magnesium + oxygen = magnesium oxide.

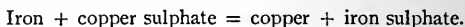
Phosphorus + oxygen = oxide of phosphorus.

Hydrogen + oxygen = water.

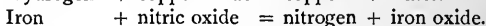
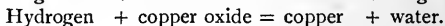
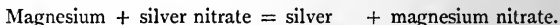
In contrast to analysis a change of this sort is called a *synthesis*.

**Decomposition and Simultaneous Combination.**—We have had frequent instances of more complicated changes than simple decompositions or combinations. Thus, when iron

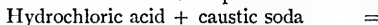
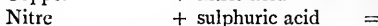
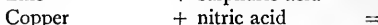
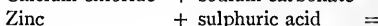
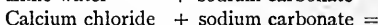
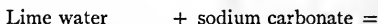
was added to a solution of copper sulphate, the copper sulphate was decomposed, copper being precipitated, but at the same time iron was dissolved and iron sulphate formed. Or,



Other instances of decomposition associated with combination were :—



Many other changes were only partially studied, and cannot as yet be classified with certainty. Try, for instance, how far you can complete the following equations as statements of the whole of the change effected :—



Although much time has been spent, in the preceding pages, in examining these and other changes, yet in very few cases has the whole course of the change been verified, and it is therefore premature to classify them definitely.

All the varied forms of change instanced above are what are called *chemical changes*, in distinction from those changes of *state*, of solid into liquid, or liquid into vapour, and the reverse, which have already been considered. Some of the characteristics of chemical change will be brought out in the succeeding pages.

The changes which matter can undergo are thus of the most varied character, and though matter may become invisible, formless, and vague, yet we assured ourselves that *matter was not destroyed* in several test cases (*see* p. 194). We therefore assumed again and again in our experiments that the sum of the weights of the substances formed as the result of a change was equal to the sum of the weights of the substances taken. As Lavoisier said, "C'est sur ce principe qu'est fondé tout l'art des



expériences en Chimie"—the whole art of making experiments in Chemistry is founded upon this principle.

### Homogeneous and Heterogeneous Substances.—

All substances may be classified into two divisions, as homogeneous or heterogeneous, according as any portion appears or does not appear, so far as our powers of observation go, to be of the same kind throughout, however minutely it is examined.

(a) *Homogeneous*.<sup>1</sup>—Such are oxygen, chalk, water, solder, a solution of salt in water, air, &c. These will be considered in detail directly.

(b) *Heterogeneous*.<sup>2</sup>—For example, a plum pudding, blood, or a mixture of sand and nitre. Plum pudding is compounded of suet, raisins, currants, lemon-peel, flour, &c., in proportions which may be varied at the discretion of the cook, but some at least of the constituents—the currants, the raisins, and the lemon-peel—may be distinguished and picked out from the finished pudding. If a drop of blood is examined microscopically, it is seen to contain minute solid discs or corpuscles of two kinds floating in a clear liquid, the one pale red in colour, the other colourless or white. So too in a mixture of sand and nitre, however intimately they are mixed together, yellow grains of sand and tiny white crystals of nitre can be readily distinguished when the mixture is examined under a microscope. If a little of the mixture is placed in the mouth the nitre is dissolved with its characteristic taste, and the hard gritty insoluble sand is left upon the tongue. It is evident that the sand and nitre are still present as such.<sup>3</sup>

Chemistry deals as a rule only with homogeneous substances; if, however, a chemist has to deal with a heterogeneous substance, he considers its homogeneous constituents separately.

Homogeneous substances can themselves be classified under three heads as—

- (1) Elements.
- (2) Chemical compounds.
- (3) Indefinite compounds.

<sup>1</sup> Greek : *homos*, the same, and *genos*, kind.

<sup>2</sup> Greek : *heteros*, other, different, and *genos*, kind.

<sup>3</sup> Heterogeneous substances, such as a mixture of sand and nitre, are often called "mechanical mixtures." The expression is, however, open to serious objection, as the only meaning to be attached to the term "mechanical" is *non-chemical*, and requires the definition of the term "chemical."

**1. Elements.**—Many homogeneous substances can undergo change by decomposition and be separated into different substances. Chalk will yield an equal weight of quicklime and carbonic acid gas. Carbonic acid itself can be decomposed, and we know it to contain both oxygen and carbon. Water can be decomposed into oxygen and hydrogen. How about carbon, or oxygen, or hydrogen? These substances can combine with others, but we have had no evidence of their decomposition into anything else. So far as we know the decomposition can go no further. Such a substance we have termed an *element*, and we may define the term thus :—

*An element is a substance which cannot by any known process be separated into two or more different substances.*

**2. Chemical Compounds.**—The term compound has been freely used in previous chapters of this book, though no statement has been made of the precise meaning to be attached to it. For instance, it has been stated that

Chalk is a compound of quicklime and carbonic acid,  
Magnesium oxide is a compound of magnesium and oxygen,  
Water is a compound of hydrogen and oxygen.

Let us consider these substances in detail. They all differ, in the first place, from a mixture of sand and nitre in being *homogeneous*.

Secondly, in them the *characteristic properties* of those substances of which they are composed have disappeared, and other properties are found. The unaided senses will never suggest that water is capable of separation into two invisible gases, and will find no evidence of the metal magnesium in the white infusible earth magnesia, nor of carbonic acid gas in chalk or marble. On the contrary, chalk, magnesium oxide, and water all possess distinctive properties which are markedly different from those of their respective constituents.

In the third place, a remarkable *definiteness of composition* has been found in chalk, magnesium oxide, and water :—

(1) *Chalk.*—The most careful analyses that we have made with powdered calcspar show that 100 grams of it always yield very nearly 56 grams of quicklime and 44 grams of carbonic acid (*see* p. 269). We have also combined quicklime with car-

bonic acid by different processes, and found that 56 grams of quicklime will in each case yield a white earth-like body resembling powdered calcspar, which when dried weighs 100 grams; each product is a compound of 56 grams of quicklime with 44 grams of carbonic acid. These bodies, therefore, possess the same quantitative composition while being also alike in other respects, and the same name "chalk" may be given to each. Chalk, then, is a substance which possesses this definite composition.

The more care experimenters have taken in studying this question the more nearly have their own experiments agreed with each other and with those of other experimenters. We, therefore, may conclude that the small differences which do exist between results point not so much to innate differences in composition as to *errors of experiment* which even the most experienced chemist cannot wholly avoid.<sup>1</sup> So satisfied are we of this definiteness of composition, that if a sample of "chalk" were given us, which on careful analysis gave different results, we should certainly conclude that it was not chalk. What then? Either some other substance or else impure chalk, that is to say chalk mixed with something else.

(2) *Magnesium Oxide*.—We have oxidised magnesium by two distinct processes, both by calcination in the air and through the action of nitric acid (*see* pp. 307 and 310), yet both have led to the same mean result, viz.:—

100 grams of magnesium combined with 66 grams of oxygen, and produced 166 grams of a white earth-like body. The products have, therefore, the same quantitative composition, while being also alike in other respects, and the same name "magnesium oxide" may be given to each.

Supposing a sample of metal did not after careful experiment give this result, we should pronounce it either not magnesium, or at any rate not pure magnesium.

(3) *Water*.—We studied the composition of water in detail, and arrived at these results:—

<sup>1</sup> This kind of evidence must of course be accepted cautiously, and many of the most remarkable discoveries of science have resulted from the close scrutiny of small "errors." We need only instance the discovery of Neptune, from small differences between the observed and calculated positions of Uranus, and the difference in density between different samples of "nitrogen" which has led to the discovery of the gas "argon" in the atmosphere.

2 volumes of hydrogen and 1 volume of oxygen combined and produced water.

2 volumes of hydrogen and 1 volume of oxygen combined and produced 2 volumes of steam.

water, when decomposed by the electric current, produced 2 volumes of hydrogen for every 1 volume of oxygen.

9 parts by weight of water are composed of 8 parts by weight of oxygen, and 1 part by weight of hydrogen.

The definiteness of composition is again very remarkable.

In contrast to the definiteness of composition of these three substances is the composition of heterogeneous substances such as a plum pudding, blood, or a mixture of sand and nitre. Supposing different portions of *one* particular mixture of sand and nitre were carefully analysed several times the same results would of course be obtained. Other mixtures of sand and nitre could, however, be made, each of which would be called "sand and nitre," but their composition would not be the same. Again, samples of blood differ in composition. So also puddings, each called "a plum pudding," may and almost certainly will differ in composition.

We have, then, in the course of our study met with various substances, of which chalk, magnesium oxide, and water are instances, to which the following statements apply:—

(1) They are homogeneous substances containing two or more elements.

(2) Some of their properties are not the same as those of their constituents, nor even a mean of those of their constituents.

(3) Two samples even when produced by different processes possess the same quantitative composition while being alike in other respects. The statement that there are substances of which this is true may be called the **LAW OF DEFINITE PROPORTIONS**.

Substances to which these three statements apply are called *definite or chemical compounds*.

Any change resulting in the production or decomposition of a chemical compound is a *chemical change*, and a chemical change is therefore associated with those features, which are proper to a chemical compound.

**3. Indefinite Compounds.**—There are other homogeneous substances containing more than one element which are not characterised by that definiteness of composition to which we have referred. Such are solutions, alloys, and mixed gases.

Solutions of salt in water, for example, each called “a salt solution,” may vary greatly in composition. Moreover, to a given solution we can within certain limits add either salt or water, and we shall obtain a fresh solution of salt in water, which will still be homogeneous.

Solder, again, is an alloy of tin and lead. We have not ourselves analysed different samples of solder, but if we did we should find that though homogeneous and closely alike in other respects, yet too great differences would exist between the ratios of lead to tin in different samples of solder to be due to any error of experiment. We may say that solder is an *indefinite* compound of tin and lead.<sup>1</sup>

**Pure Materials.**—It is not easy to prepare substances *pure*, that is, free from admixture with other substances. Adulteration Acts have made it easier for manufacturers to be honest, but it is still difficult to prepare pure materials, as may be seen by quoting from a recent price list :—

		lb.	oz.
Acid hydrochloric ... commercial ...	...	3 <i>d.</i>	
	pure ...	6 <i>d.</i>	
	pure for analysis ...	1 <i>s.</i>	
Soda caustic ... lump ...	...	8 <i>d.</i>	
	white stick ...	2 <i>s.</i>	3 <i>d.</i>
	pure ...	...	6 <i>d.</i>
Zinc ... granulated ...	...	6 <i>d.</i>	
	free from arsenic ...	2 <i>s.</i>	
	distilled, free from iron	6 <i>s.</i>	

Each of the above substances is homogeneous : the first is an indefinite compound, the second a definite or chemical compound,

<sup>1</sup> We do not contend that our treatment of chemical and indefinite compounds is complete or above criticism. On the contrary, we believe that at present it is impossible to define the one clearly from the other. This is shown by the fact that there are some substances with regard to the classification of which as chemical or indefinite compounds there is wide difference of opinion amongst even trained chemists.

and the third an element. It is clear, however, that the commercial article may in each case contain foreign matter.

Boyle, in *The First Essay on the Unsuccessfulness of Experiments*, speaks of seed of clover grass from a great quantity of which no grass sprang up, and of varying samples of vinegar which did not all contain living creatures, and then of his own difficulties, thus :—

“The other day, having occasion to use some *Spirit of Salt*, whereof I was not then provided, I sent for some to a Chymist, who making it himself, was the likelier to afford that which was well made : but though I gave him his own rate for it, at the first Rectification even in a Retort, a single pound afforded us no less than six ounces of phlegm ; and afterwards being further rectif’d in a high body and gentle heat, the remaining Spirit parted with a scarce credible quantity of the like nauseous liquor, and after all these sequestrations of phlegm was not pure enough to perform what we expected from it. Of which complaining to an excellent Chymist of my acquaintance, he sent for Spirit of Salt to a very eminent distiller of it, who gets much by his profession, and passeth for a very honest man : but this spirit, besides its weakness, discover’d itself to be sophisticated with either Spirit of *Nitre*, or *Aqua fortis*, which betray’d itself by its peculiar and odious smell ; whereas Spirit of Salt skilfully and sincerely drawn, is commonly of a greenish colour, bordering upon yellow, and hath usually a Peculiar, and sometimes (as I can exemplifie to you in some of mine) a not Unpleasing smell.”

The greenish colour which Boyle regarded as a test of purity we should now think a proof of impurity due to the presence of iron.

**The Existence of an Element in the Separate and in the Combined State.**—We have stated that magnesium oxide is a chemical compound of magnesium and oxygen. We do not mean, however, that the white infusible earth actually contains magnesium as a *metal*, and oxygen as a *gas*. That is inconceivable. What we do mean is that magnesium oxide contains the material or substance of these elements. A similar statement is true of other chemical compounds. Water, for

instance, does not contain oxygen or hydrogen as gases, but the substance of these elements. Why, when elements are thus intimately compounded, their characteristic properties should vanish, and why other properties should appear, is, at present, an unanswered question.

Hence, to be accurate, a statement of the composition of a chemical compound is merely the statement of its formation or of the changes of which it is capable.

## PRACTICAL EXAMINATIONS

THE following papers will require about two hours each. The answers should be sent up in writing, and should show clearly—(1) what the student does, (2) what he observes, (3) what conclusions he comes to.

It is desirable for the student to plan out his work so that experiments requiring time, *e.g.*, evaporations and the preparation of crystals, may be started at once, and shorter exercises worked out in the intervals of longer ones.

### I

1. Make a list of, and add up the weights on the balance pan.
2. Find the area of the post-card, and express the answer (*a*) in sq. inches, (*b*) in sq. c.c.
3. Read the level of the liquid in the burette.
4. Use the 100 c.c. flask to find the relative density of alcohol.
5. Find the percentage of water of crystallisation in A.

### II

1. Place the two scales edge to edge, and read off the number of millimetres in 1, 2, 3, 4, 5, and 6 inches respectively.
2. Read the temperature shown on the thermometer standing in the flask of water.
3. Find the relative density of the brass chain.
4. From the given blue litmus solution prepare a red litmus solution. Prepare also solutions of three different tints, neither red nor blue, but intermediate between these colours.
5. Find the percentage of sand in the given mixture of sand and nitre



## III

1. Weigh the given coin to the nearest centigram.
2. Given a 100-grain pipette, and a burette graduated in c.c., find the number of grains in 1 c.c.
3. Make an egg-blowpipe from the glass tubing.
4. Find the melting point of the substance A.
5. Describe the effect of heat on the substance B.

## IV

1. Criticise the apparatus arranged for preparing and drying carbonic acid gas.
2. Find the boiling point of the liquid A.
3. Describe the appearance of the substance B.
4. Find whether C, D, and E are soluble in water or not.
5. F is a solution of an acid, and G one of an alkali. Prepare from them clean and dry crystals of a salt.

## V

1. A, B, and C are three different acids. Apply tests, and find out what they are.
2. Evaporate to dryness 50 c.c. of the liquid D, and find the number of grams of solid material dissolved in 1000 c.c.
3. Describe the effect of heat on the substance E.
4. Read the vernier on the barometer.
5. Prepare a sample of crystallised bichromate of potash from the given mixture of chalk and bichromate of potash.

## VI

1. Find whether any of the substances A, B, C, and D liberate Fixed Air (carbonic acid gas) in contact with hydrochloric acid.
2. Name the specimens E, F, G, and H.
3. Find whether the water in the bottle J is hard or soft.
4. Find what gas is contained by the corked flask K.
5. Make a careful drawing of a crystal of quartz.

## VII

1. Make a list of and add up the weights missing from their places in the box.
2. Read the level of the liquid in the measuring jar.
3. A chalk mark extends along part of the length of the given scale. Express the chalked length as a decimal of a metre.
4. Weigh an empty beaker. Measure into it ten successive quantities of 100 grains each from a 100-grain pipette. Weigh again. Find in grams the weight of 100 grains of water.

5. Prepare from zinc some crystals of zinc sulphate free from zinc and free from excess of sulphuric acid.

6. A gas is given off when strong nitric acid is added a little at a time to a cold saturated solution of ferrous sulphate and the mixture is gently warmed. Collect some of the gas and examine its properties.

### VIII

1. Act on the metal A with the three acids—sulphuric, hydrochloric, and nitric—which are provided. Collect the gases evolved, if any, and study their properties.

2. Prepare if possible crystallised samples of the salts formed.

### IX

1. Find whether the substances A, B, C, and D give off Fire-Air (oxygen) when heated in a test-tube.

2. The Solution E contains ten grams of sulphuric acid per litre. Find the number of grams of acid in one litre of F.

3. Find some solvent for the substance G.

4. Find what weight of the metal H will replace 100 grams of copper.

### X

1. Describe the effect of heat on the substance A.

2. Find what weight of carbonic acid can be obtained from 1 gram of B.

3. Oxidise the metal C with nitric acid, and find what weight of oxygen will combine with 100 grams of it.

### XI

1. Find how many c.c. of the alkali A are equivalent to 1 c.c. of the acid B.

2. You are provided with a sample of litharge. Convert it into lead carbonate by any method you can devise. Experiment first with small quantities, then with a weighed quantity. Collect the lead carbonate carefully on a filter-paper, wash, and dry at  $100^{\circ}$ .

### XII

1. Find the effect of passing a stream of hydrogen over the oxides A, B, and C, in a glass tube heated by a Ramsay burner.

2. Find the volume of hydrogen evolved when one to two grams of the zinc dust provided are dissolved in hydrochloric acid. Calculate (a) the volume of hydrogen obtained from one gram of the zinc dust, (b) the percentage of metallic zinc in the dust.

## APPENDIX

### LABORATORY FITTINGS.—APPARATUS AND CHEMICALS

**The Laboratory**,<sup>1</sup> in which the student is to work will probably be found by him ready fitted. If he has to fit up his own, the most important requirements are :—

- A room of sufficient height, well lighted and ventilated.
- A bench of some hard wood, such as teak, not easily destroyed.
- Shelves for bottles above the bench.
- Drawers and cupboards for apparatus, below the bench.
- Gas for heating and lighting.
- A water supply.
- An earthenware or lead-lined sink, not easily choked by rubbish.
- A draught-cupboard for substances evolving noxious fumes.

**Plans for Laboratory Fittings** may be obtained from :—

- North of England School Furnishing Co., Darlington ;
- Hammer and Co., London ;
- John Heywood, Deansgate, Manchester ;
- The Science and Art Department's Drawings and Descriptions of Fittings for a Laboratory (Form No. 1013), Eyre and Spottiswoode ; 1s. 6d.

#### APPARATUS FOR EACH BENCH.

Each student should have, for his own individual use, some such set of apparatus as the following :—

- Retort-stand with two rings.
- One clamp for do.

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<sup>1</sup> Latin, *laborare*, to work.

Iron tray for sand, or piece of asbestos board.  
Iron wire gauze.  
Three blocks of wood.  
Duster.  
Test-tube stand.  
Twelve test-tubes ;  $6 \times \frac{3}{4}$  inch.  
Test-tube brush. ✓  
Wooden scale (inch and centimetre).  
Watch glass.  
Three flasks ; 16 oz., 8 oz., and 4 oz.  
Two beakers ; No. 3 and No. 4.  
Two porcelain evaporating basins.  
Two funnels.  
Filter-papers.  
Thistle-funnel.  
Two porcelain crucibles and lids.  
Pipe-clay triangle.  
Tripod-stand.  
Crucible tongs.  
Bunsen gas-burner and two feet rubber tubing.  
One foot narrow indiarubber tubing.  
Mohr's spring clip.

#### APPARATUS FOR GENERAL USE

In addition, the following apparatus will be frequently wanted, and must be close at hand, and in sufficient quantity.

Flasks ; various sizes (20, 16, 8, 4, 2 oz.).  
Beakers ; various sizes.  
Test-tubes ; various sizes.  
Woulff's bottles ; 10 oz.  
Eight oz. round-bottomed flasks, or boltheads.  
Stoppered hard glass retorts ; various sizes.  
Winchester pint bottles.  
Bell-jars.  
Soft glass tubing.  
Combustion glass tubing ;  $\frac{1}{2}$  inch internal diameter, cut into 1 foot lengths.  
Hard glass test-tubes ;  $6 \times \frac{3}{4}$  inch.  
Hard glass reduction tubes.  
Glass rods.  
Glass evaporating basins.  
Glass crystallising dishes.  
Watch and clock glasses ; various sizes,

- Calcium chloride U-tubes.
- U-tubes ;  $\frac{1}{2}$  inch internal diameter.
- One large funnel ; 8" diameter.
- Lamp glasses.
- Measuring flasks ; 1000 c.c., 250 c.c., and 100 c.c.
- Measuring jars ; 1000 c.c., 250 c.c., and 100 c.c.
- Pipettes ; 20 c.c., and 10 c.c.
- Burettes ; to hold 50 c.c., and graduated to  $\frac{1}{10}$  c.c.
- Burette-stands.
- Wedgwood-ware pestles and mortars.
- Iron pestle and mortar.
- Chemical thermometers (reading to 110° and to 300°).
- Pneumatic troughs.
- Beehive-stands.
- Gas collecting jars.
- Round plates of ground glass ; 4 inches diameter.
- Porcelain boats.
- Black indiarubber tubing,  $\frac{3}{16}$  inch.
- Pump tubing.
- Corks ; various sizes.
- Rubber stoppers ; various sizes.
- Balances ; to carry 100 grams on each pan and turn with two milligrams (one at least for every six students ; better one for every two students).
- Sets of weights ; second quality, up to 50 grams and down to one milligram.
- Salter's spring balances ; to carry 1000 grams (for rough weighing).
- Small retort-stands.
- Fletcher's radial or solid flame burners.
- Ramsay burners.
- Deflagrating spoons.
- Iron spoons.
- Cylindrical tins or cans, for boiling water ; quart size.
- Small brass and zinc cans, for calorimetry (*see* p. 95).
- Triangular and rat-tail files.
- Cork borers
- Mohr's clips
- Screw pinchcocks.
- Mouth blowpipes.
- Boxes of litmus papers.
- Platinum wire.
- Platinum foil.
- Emery paper.

## APPARATUS FOR MORE OCCASIONAL USE

Cryophorus.  
 Wet and dry bulb thermometer.  
 Grove's battery.  
 Bichromate battery and induction coil, or a small dynamo.  
 Eudiometers (U-shaped).  
 Hempel gas burettes.  
 Kipp's apparatus (two, for carbonic acid gas and hydrogen).  
 Combustion furnace.  
 Muffle furnace.  
 Foot-bellows and blowpipe.  
 Liebig's condenser.  
 Lead, iron, and copper cylinders; say 50 grams,  
 Drying ovens.  
 Platinum crucible.  
 Water-bath.  
 Balance turning with 0.2 milligrams.  
 Horse-shoe magnets.  
 Air pump.  
 Set of grain weights.  
 Exhausting air syringe  
 Water vacuum pump.

## CHEMICALS FOR GENERAL USE.

Acids—Acetic.  
     Hydrochloric, commercial and pure.  
     Nitric.  
     Oxalic, pure.  
     Sulphuric, commercial and pure.  
     Tartaric.  
 Alcohol—Methylated.  
 Alum—Ammonia.  
     Chrome.  
     Potash.  
 Aluminium, sheet.  
 Ammonia solution.  
 Ammonium—Acetate.  
     Carbonate.  
     Chloride.  
     Nitrate.  
 Antimony.  
 Asbestos.  
 Benzene.  
 Bismuth Oxide.

Borax. ✓

Calc Spar.

Calcium—Carbonate (chalk), B. P. ✓

Chloride, fused.

Oxide (quicklime), commercial.

Charcoal, wood. ✓

Chalk.

Chloroform.

Chromium Oxide.

Cobalt Oxide.

Copper—Metal, pure foil and turnings.

Carbonate.

Nitrate.

Oxide, black, pure.

Sulphate, commercial. ✓

Ether, methylated.

Flint.

Fusible metal.

Galena.

Glycerine.

Gypsum.

Iodine.

Iron—Metal, filings and pure wire.

Ammonium sulphate.

Oxide, magnetic and sesqui-

Pyrites.

Sulphate. ✓

Lead—Metal.

Carbonate.

Nitrate.

Red.

Oxide (litharge).

Peroxide.

Litmus, pure granular.

Logwood.

Magenta.

Magnesium—Metal, ribbon and wire.

Carbonate.

Nitrate.

Oxide.

Sulphate. ✓

Manganese, Dioxide.

Marble, white.

Methyl orange.

Mercury—Metal.

Iodide (red).

Nitrate.

Oxide (red).

Olive oil. ✓

Paraffin, solid white. ✓

Phenol phthalein.

Phosphorus, yellow and red.

Plaster of Paris.

Potassium—Bichromate. ✓

Carbonate.

Chlorate.

Chloride.

Ferricyanide.

Ferrocyanide.

Hydrate, sticks, pure.

Nitrate.

Permanganate.

Sulphate.

Quartz, crystals.

Rock Salt.

Sand.

Silver, Nitrate.

Sodium—Bicarbonate. ✓

Carbonate. ✓

Chloride. ✓

Hydrate, lump and sticks.

Nitrate.

Potassium Tartrate (Rochelle Salt). ✓

Sulphate.

Solder.

Starch. ✓

Sugar, cane.

✓ Sulphur, roll.

Test-papers, litmus.

Tin—Metal, granulated.

Oxide (white).

✓ Turpentine.

Zinc—Metal, granulated.

foil, pure.

Oxide.

Sulphate.



## NOTES ON APPARATUS

**Burette Taps** should periodically be greased with vaseline, and be kept in place with small rubber bands.

**Deflagrating Spoons** are cleaned by holding them in the fire for a few seconds.

**The Grove's Battery.**—When the cell is taken to pieces, the nitric acid may be kept to use again. The dilute sulphuric acid and zinc sulphate from the outer cell must generally be thrown away. The porous cells should be soaked in fresh water for a day or two with one or two changes until the water tastes quite sweet. Then allow them to dry. The zinc plates should also be soaked in water. Before fitting up the Grove cell all ends of wires and clamps must be brightened with a file, or with sand-paper, to ensure a good metallic contact.

The zinc plates may be cleaned by scraping and washing (if greasy wash with caustic soda). Dip them into dilute hydrochloric acid, and wash again. Rub some mercury over the surface of the plates. Dip them again in acid and then pour more mercury over them. Repeat this until the zinc will stand in the acid without effervescing.

To charge the battery, fill the porous cells with fuming nitric acid. Fill the outer cells two-thirds full with a mixture of one part of sulphuric acid and eight parts of water.

**Ruhmkorff Coil—Bichromate Cell.**—For a full account of a Ruhmkorff coil the student must refer to some book on electricity such as Silvanus Thompson's *Electricity and Magnetism*. For the present he may understand it as a machine for converting a steady continuous current of electricity into an intermittent series of violent sparks. The coil may be used in connection with the Grove's battery, but as this takes some time to set up it is more convenient to use a bichromate battery which can be thrown in or out of action in a moment by raising or lowering the zinc plates. The zinc plates ought to be amalgamated like those of the Grove's battery. The bichromate battery is charged with a mixture in the proportion of

Potassium bichromate	...	...	...	...	13 grm.
Water	...	...	...	...	140 c.c.
Sulphuric acid	...	...	...	...	24 c.c.

This mixture is of a clear orange colour when first prepared, but gradually changes to dark green and deposits crystals of chrome alum in the cell.

**Kipp's Apparatus** (Fig. 137) is the best for supplying a continuous current of hydrogen unmingled with air. The bulbs B and C are in one

piece. Into these is fitted the funnel A. Removing A and F, zinc rods can be slipped into B. The glass collar D is pierced with holes allow-

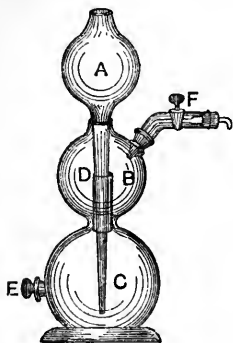


FIG. 137.

ing liquid to drain down into C, but stopping any solid material. The dilute sulphuric acid should be cooled before putting it in, as the thick glass is apt to crack. But hydrochloric acid will do at least as well, and saves this risk. To fill the apparatus open the tap F. Pour acid into A until C is full. Air is driven out through F. Close F. Pour in more acid until A is two-thirds full. Acid cannot enter B because B is full of air. Open the tap F. Air escapes under the pressure of the acid column in A. Acid enters B and effervescence begins. Close F. The effervescence continues, but the hydrogen, being unable to escape, accumulates in B. The acid is driven back into C, and up into A. The

effervescence stops as soon as the acid has drained off the zinc. The stopper E is useful when washing out the apparatus. Sulphuretted hydrogen, hydrochloric acid, and carbonic acid gas may be prepared in the same apparatus.

## ANSWERS

### EXAMPLES I.

*Metric Prefixes : Length.* (Page 9.)

- |                  |                 |
|------------------|-----------------|
| 1. 1,000,000 mm. | 10. 12345·6 cm. |
| 2. ·1 dm.        | 11. 40.         |
| 3. ·01 Dm.       | 12. 5,000.      |
| 4. ·000,001 Km.  | 13. 4 cm.       |
| 5. ·9 cm.        | 14. 2 Km.       |
| 6. 1·001 Km.     | 15. 120 cm.     |
| 7. 1 Dm.         | 16. 4.          |
| 8. 45,600 cm.    | 17. 40,000.     |
| 9. ·123 Dm.      | 18. 4.          |

### EXAMPLES II.

*Metric Square Measure : Area.* (Page 11.)

- |                    |  |
|--------------------|--|
| 1. 10,000 sq. cm.  | 9. 20 cm.  |
| 2. ·01 sq. cm.     | 10. 1 sq. dm.  |
| 3. 625 sq. mm.     | 11. $1\frac{1}{3}\frac{1}{4}$ sq. yds. ; $11\frac{1}{9}$ sq. ft. |
| 4. 30 cm.          | 1600 sq. in.   |
| 5. 10,200 sq. cm.  | 12. $196\frac{3}{7}$ cwt. ; 22,000 lbs.                          |
| 6. ·000,296 sq. m. | 13. 25,000.  |
| 7. 20 sq. cm.      | 14. 64.  |
| 8. 6 sq. cm.       |  |

### EXAMPLES III.

*Metric Cubic Measure.* (Page 12.)

- |                      |                         |
|----------------------|-------------------------|
| 1. 27,000 cc.        | 5. ·001 c.c.            |
| 2. 15,625 cu. mm.    | 6. 1000,000,000 cu. mm. |
| 3. 64 cu. ins.       | 7. ·000,001 cu. m.      |
| 4. 1,000,000 cu. mm. | 8. 1,002 cu. mm.        |

EXAMPLES III.—*Continued.*

- |                   |                     |
|-------------------|---------------------|
| 9. 3,004,000 c.c. | 14. 4320 cc.        |
| 10. 4'994 cc.     | 15. 194,400 cu. dm. |
| 11. 1331 cu. mm.  | 16. 1 cc.           |
| 12. 600 cu. mm.   | 17. 10 sq. cm.      |
| 13. 27 cu. dm.    |                     |

## EXAMPLES IV.

*Metric Capacity Measure.* (Page 13.)

- |               |                     |
|---------------|---------------------|
| 1. 1200 cc.   | 10. 1 dl.           |
| 2. 34 cc.     | 11. '3 litre.       |
| 3. 570 cc.    | 12. 3000 cc.        |
| 4. 80'6 cl.   | 13. 5.              |
| 5. 800 cc.    | 14. 2 litres.       |
| 6. '1.        | 15. 400.            |
| 7. 100 times. | 16. 324,000 litres. |
| 8. 1'234 l.   | 17. 8½ days.        |
| 9. 56,780 cc. |                     |

## EXAMPLES V.

*Metric System: Weight and Volume.* (Page 14.)

- |                |                |
|----------------|----------------|
| 1. 1150 grams. | 4. 50 grams.   |
| 2. 25 grams    | 5. 1350 grams. |
| 3. 8 kilos.    | 6. 55 cc.      |

## EXAMPLES VI.

*Metric System: Miscellaneous.* (Page 16.)

- |                   |                        |
|-------------------|------------------------|
| 1. 20.            | 6. '01 grams.          |
| 2. 50.            | 7. 500 cc.             |
| 3. 20 milligrams. | 8. 2777'7 cm. per sec. |
| 4. 1360 grams.    | 9. '0648 grams.        |
| 5. '0002 grams.   | 10. £15. 7s. 2½d.      |

## EXAMPLES VII.

*Burette-Reading.* (Page 32.)

- |           |              |
|-----------|--------------|
| 1. A. 27. | 2. I. 1'675. |
| B. 43.    | K. 13'325.   |
| C. 6'7.   | L. 10'675.   |
| D. 830.   | M. 23'25.    |
| E. 113'3. | O. 46'75.    |
| F. 136'6. | P. 632'5.    |
| G. 233'3. | Q. 867'5.    |
| H. 266'6. | R. 675.      |

## EXAMPLES VIII.

*Interpolations: Conversion of Thermometric Scales.* (Page 44.)

- |                         |              |
|-------------------------|--------------|
| 1. 29.605.              | 12. 35° C.   |
| 2. 63.                  | 13. 50° C.   |
| 3. 1.31.                | 14. 68° F.   |
| 4. 33.3.                | 15. 23° F.   |
| 5. £6 10s. od.          | 16. 40° F.   |
| 6. 5.58 $\frac{1}{2}$ . | 17. 4° C.    |
| 7. 1.361.               | 18. - 10° C. |
| 8. 59° F.               | 19. - 20° C. |
| 9. 41° F.               | 20. 12° R.   |
| 10. 176° F.             | 21. 68° F.   |
| 11. 10° C.              |              |

## EXAMPLES IX.

*Relative Densities of Liquids.* (Page 51.)

- |                   |                |
|-------------------|----------------|
| 1. 13.76; .07268. | 8. 150 cc.     |
| 2. 38.09 cc.      | 9. 250 cc.     |
| 3. 70.86 grm.     | 10. .923.      |
| 4. 36.2 grm.      | 11. 2800 grm.  |
| 5. .940.          | 12. .98.       |
| 6. 1.026.         | 13. 20 sq. mm. |
| 7. 60 cc.         |                |

## EXAMPLES X.

*Relative Densities of Solids.* (Page 57.)

- |                   |                                |
|-------------------|--------------------------------|
| 1. 2.5; 2.14; 10. | 8. 22.8 grams.                 |
| 2. (1) 8.95.      | 9. .001 cc.                    |
| (2) 7.2.          | 10. 11.4.                      |
| (3) .25.          | 11. 145 grams.                 |
| (4) 11.4.         | 12. 8.6.                       |
| 3. 17.5.          | 13. 12.5 per cent.             |
| 4. 376 grains.    | 14. 2.4.                       |
| 5. 7.7 lbs.       | 15. 179 $\frac{89}{100}$ .     |
| 6. 8.             | 16. £0 os. 11 $\frac{1}{4}$ d. |
| 7. 50 oz.         | 17. 0.8.                       |

## EXAMPLES XI.

*Specific Heat.* (Page 104.)

- |                |                   |
|----------------|-------------------|
| 1. .057 +.     | 6. 51° 2' + C.    |
| 2. .034 +.     | 7. 2° 9' + C.     |
| 3. .5.         | 8. 154.3 + grams. |
| 4. 3.75 C.     | 9. 333.3 grams.   |
| 5. 0° 94' + C. |                   |

## EXAMPLES XII.

*The Densities of Gases.* (Page 238.)

- |                 |                   |
|-----------------|-------------------|
| 1. 1'28 gram.   | 6. 1011'85 grams. |
| 2. 1'3 gram.    | 7. 1'56 gram.     |
| 3. 775'19 cc.   | 8. 60 grams.      |
| 4. 1'215 oz.    | 9. '68 gram.      |
| 5. 124'7 grams. |                   |

## EXAMPLES XIII.

*Boyle's Law.* (Page 248.)

- |                                       |                                    |
|---------------------------------------|------------------------------------|
| 1. 735'15 or 735 $\frac{15}{101}$ cc. | 7. 43'48 or 43 $\frac{11}{33}$ cc. |
| 2. 1500 cc.                           | 8. 4 ins.                          |
| 3. 742'57 or 742 $\frac{58}{101}$ cc. | 9. 60 ins.                         |
| 4. 288 cu. ins.                       | 10. 250 mm.                        |
| 5. 1200 lbs. per sq. in.              | 11. '998.                          |
| 6. 40 mm.                             | 12. 1'000.                         |

## EXAMPLES XIV

*Charles' Law.* (Page 257.)

- |                                      |              |
|--------------------------------------|--------------|
| 1. (1) -272°'97 C.                   | 3. 954'54.   |
| (2) -271°'51 C.                      | 4. 1076'92.  |
| 2. (1) $\frac{1}{272'8}$ or '003665. | 5. 1'17      |
| (2) $\frac{1}{270'13}$ or '00370     | 6. 380'0082. |
|                                      | 7. 4095° C.  |
|                                      | 8. 319° C.   |

## EXAMPLES XV.

*Charles' and Boyle's Law.* (Page 258.)

- |                   |                  |
|-------------------|------------------|
| 1. (a) 78'75 cc.  | 3. (a) 322°'67.  |
| (b) 82'11 cc.     | (b) 162°.        |
| (c) 105 cc.       | (c) 46°.         |
| 2. (a) 556'78 cc. | 4. (a) 709'8 mm. |
| (b) 560'44 cc.    | (b) 599'04 mm.   |
| (c) 476'19 cc.    | (c) 967'2 mm.    |

## EXAMPLES XVI.

*Dalton's Law of Mixed Gases.* (Page 265.)

- |              |               |
|--------------|---------------|
| 1. 700 mm.   | 5. 5013'3 cc. |
| 2. 31'9 mm.  | 6. 932'9 cc.  |
| 3. 592'5 mm. | 7. 92' 6 cc.  |
| 4. 96 mm.    | 8. 99'6 cc.   |

## EXAMPLES XVII.

*The Densities of Gases.* (Page 266.)

- |                   |                                  |
|-------------------|----------------------------------|
| 1. 1.44 grams.    | 7. (.1597 + .6111) or .7701 l.   |
| 2. .5050 litre.   | 8. (.3024 + .9954) or 1.2978 gm. |
| 3. 14.814 litres. | 9. 14.34.                        |
| 4. .5865 gram.    | 10. 16 grams.                    |
| 5. 103.9 grams.   | 11. 720 mm. and 45 mm.           |
| 6. .0264 gram.    | 12. 1391 mm.                     |

## EXAMPLES XVIII.

*Equivalents of the Metals.* (Page 314.)

1. 265.5 grams.
2. 175.1 grams.
3. (a) 226.2 grams.
- (b) 267.2 grams.

THE END.









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